TEACHER'S EDITION

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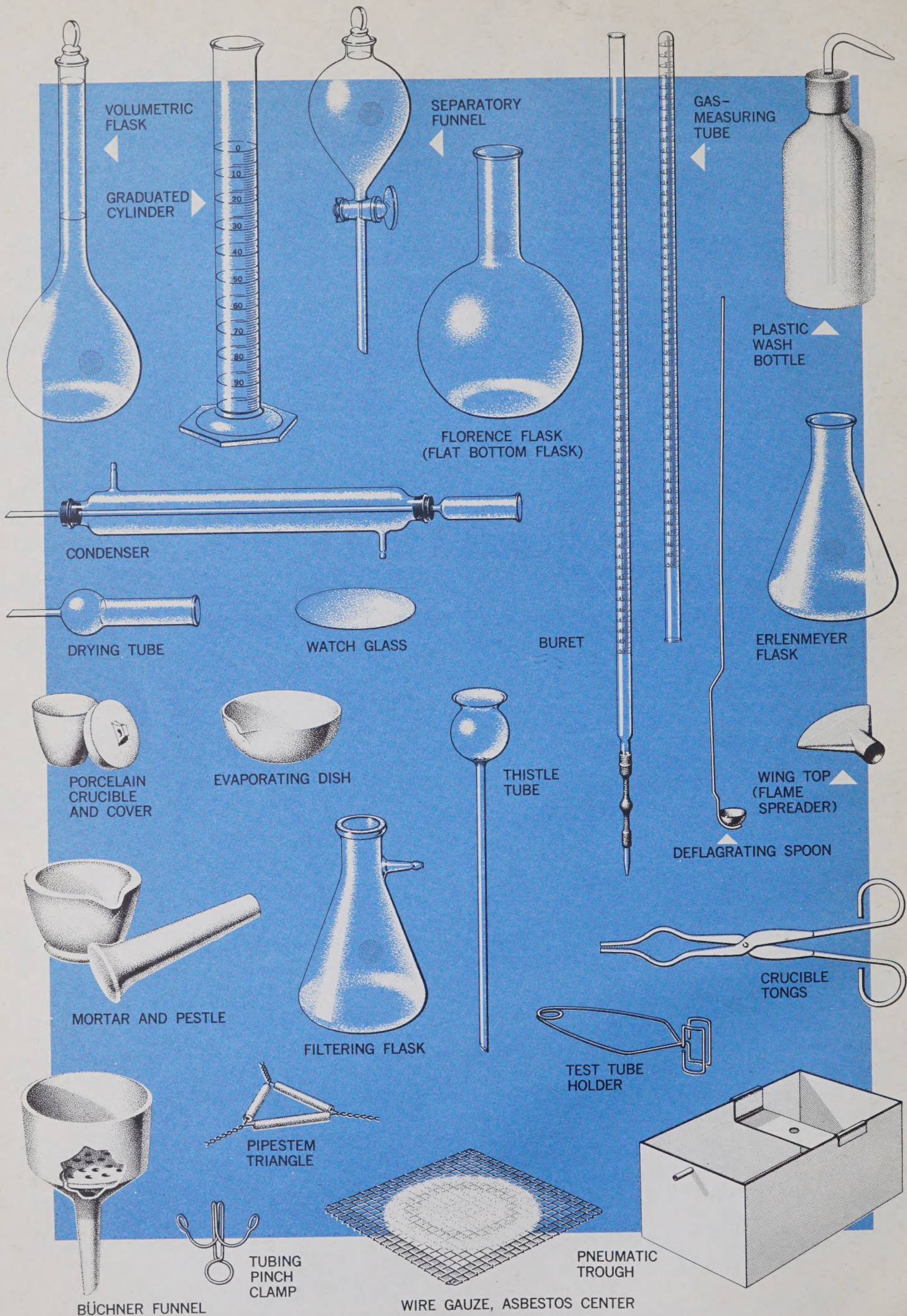
# Chemistry of Common Substances


VOLUMETRIC FLASK
GRADUATED CYLINDER
SEPARATORY FUNNEL
GAS-MEASURING TUBE
PLASTIC WASH BOTTLE
FLORENCE FLASK (FLAT BOTTOM FLASK)
CONDENSER
DRYING TUBE
WATCH GLASS
BURET
ERLENMEYER FLASK
PORCELAIN CRUCIBLE AND COVER
EVAPORATING DISH
THISTLE TUBE
WING TOP (FLAME SPREADER)
DEFLAGRATING SPOON
MORTAR AND PESTLE
FILTERING FLASK
CRUCIBLE TONGS
TEST TUBE HOLDER
PIPESTEM TRIANGLE
TUBING PINCH CLAMP
PNEUMATIC TROUGH
BÜCHNER FUNNEL
WIRE GAUZE, ASBESTOS CENTER

# Chemistry of Common Substances

**LEE R. SUMMERLIN**
Professor of Chemistry
University of Alabama in Birmingham

**TEACHER'S EDITION**

**SILVER BURDETT COMPANY**
Morristown, New Jersey
Glenview, Illinois • Palo Alto • Dallas • Atlanta

## ACKNOWLEDGMENTS

This collection of experiments has resulted from workshops, short courses, and miniterm and in-service programs over the past six years with junior-high science teachers and chemistry teachers. I would like to express my gratitude to the more than 200 science teachers who have tried the experiments, suggested changes, and recommended additional experiments. I would also like to thank the many teachers, all over the country, who have given me suggestions. Finally, I would like to personally thank my colleagues in the Chemistry Department of the University of Alabama in Birmingham for encouraging the development of this laboratory manual and supporting the programs designed to upgrade chemistry teaching at the precollege level.

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# Rules of the Laboratory

The chemistry laboratory can be a safe place in which to work. This requires, on the part of each student, a constant awareness of a few basic rules of laboratory conduct.

1. An unauthorized experiment must never be performed.
2. Safety eyeglasses or prescription eyeglasses should be worn at all times.
3. Read an assigned investigation carefully before beginning it. Note every caution listed in the experiment and be fully prepared to heed each one.
4. Never point a test tube while it is being heated, or one in which a reaction is occurring, at another student or at yourself.
5. Know what you are doing. Be wary of what neighboring students are doing.
6. Before leaving the laboratory, be certain that the gas line at your laboratory space is shut tightly. Also, check to see that all water spigots are firmly closed.

# First Aid in the Laboratory

If an accident should occur in the laboratory, there are a few basic rules to follow that can be of immediate benefit.

1. Immediately report any accident or injury, no matter how minor, to your teacher.
2. If acid splashes in your eyes, wash them out with plenty of water. Follow this by a wash with a dilute solution of sodium bicarbonate. Finally, rinse with plenty of water.
3. If base splashes in your eyes, wash them out with plenty of water. Follow this by a wash with a boric-acid solution. Rinse with plenty of water.
4. If a chemical solution other than an acid or a base splashes in your eyes, wash them out with plenty of water. Report the accident to your teacher, who may then recommend further action.
5. If acid splashes on your skin, wash the affected area with plenty of water. Apply a paste of sodium bicarbonate and water, and allow this to remain for several minutes before washing it off.
6. If base splashes on your skin, wash the affected area with plenty of water. Then wash the area with boric-acid solution. Follow this with a final wash with plenty of water.
7. For cuts and burns, follow the procedures accompanying the first-aid kit located in the laboratory.

# To The Teacher

This Laboratory Manual contains 67 experiments suitable for use in a high-school chemistry course. Some of the experiments can also be used in biology classes and junior-high science classes. All of the experiments have the following features:

1. They are easy.
2. They require relatively few chemicals and materials.
3. They are centered around common substances usually found in the home.
4. They are usually qualitative rather than quantitative. Most of the experiments require very little mathematical treatment of data.
5. They emphasize observation and recording of data and information.
6. They are short—most can be completed in 30 to 45 minutes.
7. They are FUN! They are purposely designed to be enjoyed by the student.

These experiments can serve as your primary laboratory program. The experiments may also be used to supplement a more rigorous laboratory program. Most of these experiments introduce the student to important topics in chemistry. Hopefully the student will be excited enough to seek additional information in and out of class.

The experiments in this manual have been keyed to the 1978 edition of Silver Burdett's *Chemistry*. At the beginning of each experiment there is a list of the particular chapters and excursions with which the experiment can be done.

# The Chemist in the Kitchen

This experiment will permit you to explore a few physical and chemical properties of several common substances ordinarily found in the home. You will be given four powders. For simplicity, we will call these **A, B, C,** and **D.** You will also have three liquid indicators. These will be called **Indicator I, Indicator II,** and **Indicator III.** The object of this experiment is to determine some properties of the powders by using the indicators. Do not be concerned with actually identifying the powders. You should only make observations on the reactions between the powders and indicators.

After you have determined some physical and chemical properties of the powders, you will use your observations to identify some mystery powders. These mystery powders may contain **A, B, C,** and **D** in any combination.

● This is a very good experiment with which to begin your laboratory program in chemistry. The emphasis here is on the process of reasoning from laboratory observations. The intent is not to identify the powders and indicators but merely to observe how they react.

● The powders used are:
A. Baking soda (Sodium bicarbonate)
B. Laundry starch
C. Baking powder (A mixture of several substances)
D. Table salt (Sodium chloride)
The indicators are:
I Vinegar (Acetic acid)
II Iodine solution ($I_2$ and KI in alcohol)
III Distilled water

## PROBLEM

To determine some physical and chemical properties of several powders.

### Materials

Powders **(A, B, C, D)**
Indicators **(I, II, III)**
Spot plate

● Prepare the powders and indicators in small containers. If you make 10 classroom sets of powders and indicators they can be used many times.

## PROCEDURE

### Part A: Preparing a Spot Plate

If you have spot plates you may use them. Otherwise, prepare your spot plates as follows.

● For extensions of this experiment you might try other powders such as powdered milk, confectionary sugar, flour, table sugar.

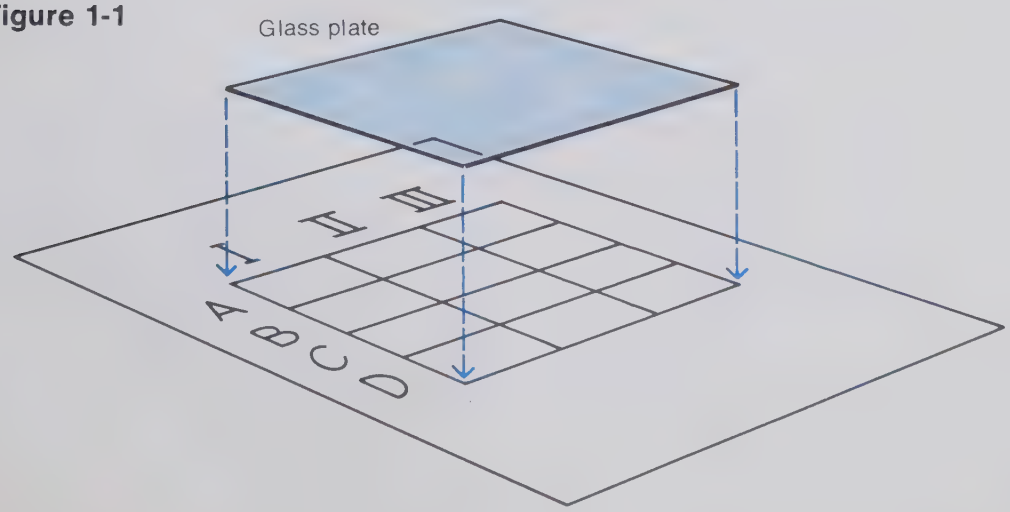
1. Place a large, glass plate over a sheet of white paper.
2. Using a pencil, outline the glass plate on the paper.

● Glass plates 6 inches on a side will work well.

3. Remove the glass plate and divide the square into 12 smaller squares by drawing three horizontal and two vertical lines on the square. (See Fig. 1-1)
4. Label the vertical columns I, II, and III. Label the horizontal columns A, B, C, and D. Replace the glass plate over the square. Now, each small square will show the reaction between a specific powder and a specific indicator.

● You may wish to make one chart and mimeograph it for the class.

**Figure 1-1**



Take this opportunity to advise students not to taste chemicals. Not only is it dangerous, but it will also give away the identity of the chemicals used in this experiment.

### Part B: Determining Physical Properties of the Powders

1. Place a pea-sized sample of each powder in the appropriate squares.
2. Examine each powder carefully. Record your observations in Table 1-1.

### Part C: Determining Chemical Properties of the Powders

1. Place 1 drop of **Indicator I** on each powder in the first vertical column.
2. Carefully note any reaction and record your observations in Table 1-2.
3. Add **Indicator II** to the powders in the second column and **Indicator III** to the powders in the third column. Record your observations.

The emphasis here is on following directions and making observations.

**TABLE 1-1 • Physical Properties**

| POWDERS | OBSERVATIONS |
|---|---|
| A | *Small white grains* |
| B | *White powder* |
| C | *White powder* |
| D | *Crystals* |

**TABLE 1-2 • Chemical Properties**

| POWDER | INDICATOR | OBSERVATIONS |
|---|---|---|
| A | I | *Foams and bubbles* |
| A | II | *Absorbs reddish color* |
| A | III | *Granular suspension* |
| B | I | *Milky mixture* |
| B | II | *Turns black* |
| B | III | *Milky mixture* |
| C | I | *Foams and bubbles* |
| C | II | *Turns black* |
| C | III | *Milky mixture* |
| D | I | *No reaction* |
| D | II | *Slight reddish color* |
| D | III | *Begins to dissolve* |

## Part D: Analyzing Mystery Powders

Now you can use the observations made with the powders and indicators to play detective. There are four mystery powders containing two powders and four mystery powders containing three powders. None of the mixtures are the same. Working with a partner, organize the work so that the composition of each mystery powder is determined. Some mixtures may be easy to analyze. Some may be difficult or perhaps impossible using only the indicators provided. Record all your observations in Table 1-3.

Perhaps you might wish to make a mixture from the known powders which you think is identical to a powder being analyzed. You can compare the known and unknown powders.

● These mystery powders should be made up and dispensed in small bottles labeled in some manner. Remember to keep a key.

● The importance of organizing the task and teamwork are emphasized here.

● Emphasize the experimental nature of chemistry—things do not always go well.

**TABLE 1-3 • Mystery Powder Analysis**

| POWDER | OBSERVATIONS | ANALYSIS |
|---|---|---|
| 1 | | |
| 2 | | |
| 3 | | |
| 4 | | |
| 5 | | |
| 6 | | |
| 7 | | |
| 8 | | |

● As an optional exercise, prepare another mystery powder which can be any one of the four powders or a mixture of any of them.

## QUESTIONS

**1.** Did you agree with all of your classmates on the results of your observations and tests?

*Answers will vary.*

**2.** If observations and results are not identical, how could you determine whose data are correct?

*Repeat the experiments, compare results further, and so on.*

**3.** Your observations in this experiment are limited to the senses. How could you extend your ability to make observations?

*Use equipment such as thermometers, balances, and so on.*

This experiment is designed to give students experience in making observations and in examining physical and chemical properties.

EXPERIMENT 2



# Investigating Milk

When a chemist analyzes fresh milk, the composition is typically: 87% water, 4.8% carbohydrate, 4.0% fat, 3.5% protein, 0.7% minerals. The minerals include calcium and phosphorus. You will test for these in a later experiment. Milk also has many vitamins, including extra vitamin D that is added at the dairy. When heated, milk reacts with acid substances to coagulate milk protein. Sour milk is a common example of milk with coagulated protein. Custards contain egg protein along with protein dissolved in milk that has been coagulated by heat. Perhaps you can think of other examples of products containing coagulated milk protein. In this experiment you will look at some of the properties of milk. You will also compare fresh and sour milk.

## PROBLEM

To examine some physical and chemical properties of milk.

Milk from the cafeteria can be used. Place one container of milk in a warm place a day or so prior to the experiment and it will be sour when you need it for this experiment.

### Materials

| | |
|---|---|
| Fresh milk | Microscope and microscope slides |
| Sour milk | Droppers |
| Vinegar | Universal pH paper |
| Evaporating dish | Small marble |
| Hotplate | Feather, hair, or wool |
| Test tubes | Stopwatch (or watch with second hand) |

## PROCEDURE

### Part A: Detecting Protein in Milk

1. First, burn a feather, some hair, or a small piece of wool. Note the unusual odor of burning protein.
2. Now, heat 3 mL of fresh milk in an evaporating dish until you can detect an odor. How does the odor of burnt milk compare with that of burnt protein?

*The odors will be similar.*

The American National Metric Council has designated L as the symbol for liter.

### Part B: Comparing Fresh and Sour Milk

1. Place 10 mL of fresh milk in a test tube. Add 10 mL of sour milk to another test tube. Describe the odor in each tube of milk.

2. Using the pH test paper, determine and record the pH of both fresh milk and sour milk.

pH of fresh milk *About 6.5* pH of sour milk *About 4.5*

At this point, explain to students that pH measures acidity, one of the chemical properties of a substance.

3. To test the thickness of milk, drop a marble into a test tube ¾ filled with fresh milk . How long does it take the marble to reach the bottom? Repeat the test with a test tube ¾ filled with sour milk. Record your observations.

Time in fresh milk Time in sour milk

● The marble falls more quickly in fresh milk.

### Part C: Microscopic Examination of Fat

● Borrow microscopes from the biology department.

1. Place 1 drop of fresh milk on a microscope slide. Add 5 drops of water and mix with a toothpick. Examine the slide under the microscope. The tiny beads are fat globules. These were formed when the milk was homogenized.
2. Repeat the experiment with 1 drop of sour milk. Describe the difference between the fresh and sour milk.

*There are no fat globules in the sour milk.*

### Part D: Curdling Milk

● You might point out that blood "curds" much like milk, producing a clot and a clear fluid, the serum.

1. Place 15 mL of fresh milk in a test tube. Add 15 mL of vinegar to the milk. Cover the top of the tube with your thumb and shake the tube vigorously. Allow the milk to settle for 5 minutes. What has happened to the milk?

*A "curd" forms. The clear liquid left is the "whey."*

2. Repeat the experiment with 15 mL of sour milk. Record your observations.

*No change*

## QUESTIONS

1. What food products are based upon the souring of milk?

*Buttermilk, sour cream, cheese, yogurt, and others*

2. What other physical properties of milk could you easily determine?

*Density, turbidity, color, and so on*

Sweetness is a property that one normally associates with sugar. But, there are several types of sugars. Are all sugars sweet? This experiment emphasizes physical and chemical properties as distinctive features of a substance.

EXPERIMENT 3

Chapters 1, 20

# Are All Sugars Sweet?

In this experiment you will compare the sweetness of several different sugars by tasting them. You will also compare the sweetness of these sugars with a sugar substitute. Since tasting chemicals can be dangerous you should never do so unless instructed by your teacher.

Emphasize to students that these are sugars, not laboratory chemicals. Thus, tasting is permitted only because you have so instructed them.

## PROBLEM

To compare the sweetness of several sugars.

### Materials

Glucose, Fructose, Sucrose, Sugar substitute
Galactose, Maltose, Lactose

Fructose is the sweetest sugar known. You might like to discuss sweeteners here. These are not sugars, but synthetic compounds that have a sweet taste.

## PROCEDURE

Place a small amount of one of the sugars in your mouth. In Table 3-1 record the relative sweetness of the sugar as very sweet, medium sweet, or not sweet. Rinse your mouth with water and taste a second sugar. Taste and record the relative sweetness of each sugar. Rinse your mouth with water between each taste test.

This is also a good opportunity to point out problems in collecting data. Since tastes vary from student to student, their answers will vary also.

**TABLE 3-1 • Relative Sweetness of Sugars**

| COMPOUND | FORMULA | VERY SWEET | MEDIUM SWEET | NOT SWEET |
|---|---|---|---|---|
| Sucrose | $C_{12}H_{22}O_{11}$ | x | | |
| Maltose | $C_{12}H_{22}O_{11}$ | | | x |
| Lactose | $C_{12}H_{22}O_{11}$ | | | x |
| Glucose | $C_6H_{12}O_6$ | | x | |
| Galactose | $C_6H_{12}O_6$ | | | x |
| Fructose | $C_6H_{12}O_6$ | x | | |
| Substitute | — | x | | |

Even though many sugars have the same molecular formulas, they have different structural formulas because the atoms are arranged differently.

*Relative Sweetness of Carbohydrates* (based on sucrose 100)

| | |
|---|---|
| Lactose | 16 |
| Galactose | 32 |
| Glucose | 74 |
| Sucrose | 100 |
| Fructose | 173 |
| Cyclamate | 3,000 |
| Saccharine | 45,000 |

## QUESTIONS

**1.** Rank the compounds in order of sweetness.

**2.** Use the library to find other differences in properties of the sugars.

**EXPERIMENT 4** Chapters 1, 2

# A Chemical Apple Pie

Our senses are very limited and easily fooled. A chemist can prepare a synthetic flavor that our senses cannot differentiate from the real thing. In this experiment, you will make an apple pie. It will be an unusual apple pie because it will contain no apples! It will, in fact, be a chemical apple pie. You will give your pie all of the characteristics of an apple pie, adding crackers to simulate the bulk of apples. You can do this experiment at home, in the home-economics laboratory, or anywhere there is an oven available. Have fun and enjoy your "apple" pie!

## PROBLEM

To demonstrate the use of artificial flavors and the limitation of the senses, and to prepare a chemical pie.

● The prepared, frozen pie crusts work well.

### Materials

● Encourage students to bring these materials from home. Round, club crackers work well.

Pastry for two pie crusts
Sugar (sucrose, $C_{12}H_{22}O_{11}$)
Cinnamon
Butter
Cream of tartar (potassium bitartrate)
Crackers (round, snack crackers)
Large pot
Pie pan

## PROCEDURE

● You will need a heat source to prepare the pie filling.

1. Bring 500 mL of water to a boil in a large pot.
2. Mix 1.5 cups of sugar with 1.5 teaspoons of cream of tartar.
3. Add this mixture to the boiling water.
4. Add 25 whole crackers, one at a time, to the water.
5. Boil for 3 minutes but **do not stir.**
6. Pour this mixture into a pastry-lined pie pan.
7. Sprinkle with cinnamon.
8. Dot lightly with butter.
9. Cover with a pastry top. Pierce the top several times with a knife to allow steam to escape.
10. Bake at 450°F for 30 minutes, or until done.
11. Cool and enjoy the results of your experiment!

● We are purposely using "cup" and "teaspoon" here to tie in chemistry with the student's everyday experience. You might wish to take the opportunity to convert to the metric system. (1 cup = 250 mL; 1 teaspoon = 5 mL, and so on.)

● The formula for potassium bitartrate is $KHC_4H_4O_6$.

## QUESTIONS

1. What flavors can you detect in your apple pie?

*Students should recognize an apple taste: possibly cinnamon, also.*

2. List some other food products which contain artificial flavors.

*Answers will vary.*

# Spinach, Catsup, and Chromatography

Chromatography is a general term for separating a substance into its component parts. *Chromos* is from the Greek word for color. Using chromotography we can separate ink into its various colored pigments. We can also separate the amino acids that occur in fruits and vegetables.

Chromatography is a very important technique in analytical chemistry. It provides a way to separate and identify components of a mixture. There are several chromatography techniques. Some are complex and require expensive materials and equipment. One simple, easy to use method of analysis is **paper chromatography.**

● Other types of chromatography include gas chromatography and liquid chromatography. You might want students to do some reading on these techniques.

The separation technique in paper chromatography depends on the migration of a mixture through the paper. Different components of a mixture will be deposited at different levels, depending upon mass, solubility, and other factors. In some cases, it is necessary to stain or develop the paper after migration has occurred.

The distance a substance travels during chromatography is given as the *rate of flow,* $R_f$. This is the ratio of the distance traveled by the substance to the distance traveled by the solvent.

$$R_f = \frac{\text{Distance traveled by substance}}{\text{Distance traveled by solvent}}$$

$R_f$ values are characteristic of a substance and can be used to identify an unknown substance. Many things can affect the rate of flow—temperature, the amount of sample used, the nature of the solvent, and so on. In this experiment you will use paper chromatography to separate the pigments ordinarily found in two common substances, spinach and tomato catsup.

## PROBLEM

To separate the pigments in tomato catsup and spinach by paper chromatography.

### Materials

● A small amount of catsup is needed—obtain it from the school cafeteria.

● This will be the first time students have used organic solvents. You must use a hood or a well ventilated area. Caution students not to place these solvents in their mouth or to breathe the fumes excessively.

Fresh spinach
Tomato catsup
Acetone
Ethyl alcohol
Petroleum ether
Beaker
Evaporating dish
Stirring rod
Wide-mouth jar or test tubes for chromatography chambers
Stopper for jar or test tube
Small capillary tubes
Filter paper or chromatography paper

## PROCEDURE

### Part A: Preparing the Paper and Chromatography Chamber

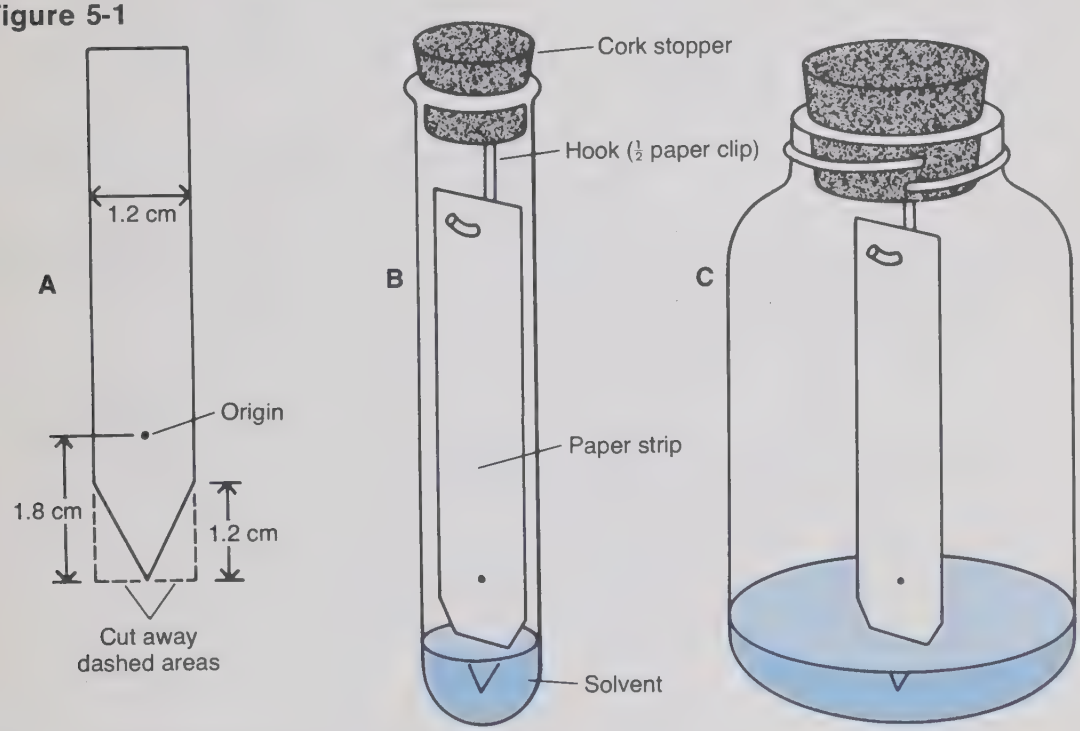
1. You can use a jar or wide-mouth test tube for the chromatography chamber.
2. Cut several strips of filter or chromatography paper about 1.2 cm wide. (The length of the paper will depend on the height of the jar or test tube you use.)
3. Cut a 1.2 cm point at one end of each paper as shown in Figure 5-1.
4. Make a pencil mark about 1.8 cm from the tip. This is the origin.
5. The paper strip can be secured to the stopper with a paper clip. Adjust the paper on the paper clip so that the tip is just above the bottom of the jar.

● Keep flames out of the lab while performing this experiment.

● Large baby food bottles are excellent for this experiment. Toothpicks can be used to spot the material on the paper.

● Be sure *not* to use a pen.

**Figure 5-1**




### Part B: The Pigments in Spinach

1. Remove the stems and lower portions of the leaves of fresh spinach.
2. Chop or cut the remaining leaves into fine pieces.
3. Weigh out about 3 g of the spinach.
4. Place the spinach in a small beaker and add 10 mL of acetone. Using a stirring rod, crush the spinach and stir the mixture for 3 to 4 minutes.
5. Pour off the acetone solution into a clean beaker and set this aside under a hood or in a well-ventilated area.
6. Add 10 mL of ethyl alcohol to the spinach. Again, crush and stir for 3 to 4 minutes.
7. Pour off the ethyl alcohol solution into the beaker containing the acetone.
8. Cover the solution in the beaker with a plastic wrap. This solution contains the extracted pigments.
9. Add solvent supplied by your teacher to just cover the bottom of your chromatography chamber. Stopper the chamber. (Be sure the paper strip is not in place.)

● Prepare the solvent in a large flask as follows:
90 mL petroleum ether
2 mL benzene
10 mL acetone
Stopper the solution.

10. Using a small capillary tube, place 1 drop of the spinach extract on the origin of your paper strip. Allow the spot to dry. Add 4 more drops, allowing each spot to dry before adding the next drop.
11. Attach the strip to the stopper and place the stopper in the chamber. The tip of the paper should just be in contact with the solvent.
12. Allow the solvent to ascend the paper until it is about 1.5 cm from the top.
13. Remove the paper and hang it up to dry. Mark the place where the solvent stopped.
14. After drying, observe the paper. How many pigments did you separate?

*Yellow, dark green, light green, possibly red-pink*

● Carotene is yellow. Chlorophyll a is dark green. Chlorophyll b is light green. Xanthrophyll is red-pink. Xanthrophyll may not separate in some samples. Catsup also has the pigment lycopene. This is the red pigment which should remain behind.

### Part C: The Pigments in Catsup

1. Place 2 mL of catsup, 2 mL of acetone, and 8 mL of petroleum ether in a test tube.
2. With your thumb over the mouth of the tube, shake the test tube.
3. Pour the liquid into an evaporating dish. Allow the clear liquid to evaporate under a hood.
4. Use the remaining liquid to spot another piece of paper as you did with the spinach extract.
5. Add fresh solvent to your chromatography chamber and hang the strip of paper in the chamber. Follow the same procedure that you used with the spinach extract.
6. After drying, observe the paper. How many pigments did you separate?

*Yellow, dark green, light green, possibly red-pink and a red pigment remaining behind.*

● The types of pigments usually found in plants are:
1. The flavonoids—red, blue, and yellow colors.
2. The carotenoids — red, orange, and yellow colors.
3. The chlorophylls—green colors.

## QUESTIONS

1. What practical applications can you suggest for chromatography?

*Identifying components of a mixture*

2. What other separation techniques are available?

*Distillation should be mentioned here. You might wish to follow this experiment with Experiment 7, Fermentation and Apple Cider.*

3. Measure the distance traveled and $R_f$ value for each pigment in both extractions.

*Answers will vary.*

EXPERIMENT 6

Chapter 1

# The Mineral Content of Milk

● The purpose of this experiment is to give students experience with analytical techniques. Be sure to demonstrate the proper way to carry and heat a crucible.

Part of milk's nutritional value is due to the minerals that it contains. In this experiment, you will determine the amount of minerals dissolved in milk. You will have available milk produced by several processes and in several forms: whole milk, skim milk, condensed milk, and powdered milk. Your teacher will indicate which milk sample you are to use in the experiment.

## PROBLEM

To determine the mineral content of various types of milk by gravitational analysis.

### Materials

Whole milk
Skim milk
Condensed milk
Powdered milk (reconstituted)
Goat's milk (if available)

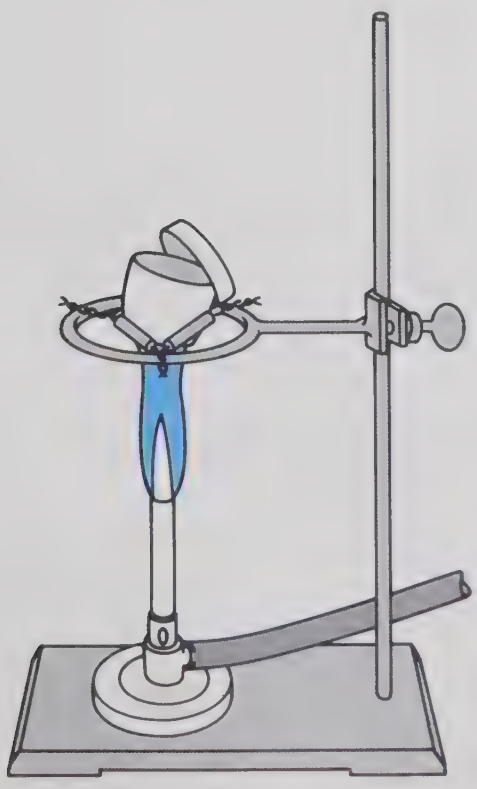
Crucible with cover
Crucible tongs
Bunsen burners
Ringstand
Clay triangle or wire screen

## PROCEDURE

In this experiment, you will determine the percentage of minerals in milk by gravitational analysis. First, you must prepare the crucible which you will use in the experiment.

1. Place a clean, dry crucible and lid on a wire screen or clay triangle supported by a ringstand. (See Fig. 6-1) Heat it strongly for five minutes. The lid should not completely cover the crucible.

**Figure 6-1**



● Students often forget to weigh the crucible cover.

2. Using the tongs, transfer the crucible and lid to a wire screen, or preferably to a dessicator to cool.
3. Transfer the crucible and lid to a balance with the tongs and determine the mass. Record the mass in Table 6-1.
4. Fill the crucible 1/3 full of the milk sample assigned to you.
5. Determine the mass of the crucible, lid, and milk and record the mass.
6. Heat the crucible *slowly* until all of the milk has evaporated. The crucible top should not completely cover the crucible while it is being heated.

● This will require intense heat. Show students how to adjust burners.

7. The milk will first char, becoming brown. Continue to heat it until it becomes a black ash. At this point, strong heating will change the black material to a white ash.
8. Cool the crucible and lid as you did before.
9. When the crucible is cool, transfer it to the balance with the tongs and determine the mass of the crucible, lid, and ash. Record your data.
10. Calculate the percent of mineral matter in the milk.

$$\frac{\text{Mass of ash}}{\text{Mass of milk}} \times 100 = \quad \% \text{ Minerals}$$

11. Compare your data with those of your classmates.

**TABLE 6-1 • Mineral Composition of Milk**

| | |
|---|---|
| Mass of Crucible and Lid | |
| Mass of Crucible, Lid, and Milk | |
| Mass of Milk | |
| Mass of Crucible, Lid, and Ash | |
| Mass of Ash | |
| Percentage of Mineral Matter | |

## QUESTIONS

● The mineral content of various types of milk follows:
Whole milk—13%
Skim milk—10%
Evap. milk—26%
Buttermilk—10%

1. How does the amount of mineral content vary with the different milk samples?

2. What experimental errors would cause mineral percentage to be greater than it should?

*Failure to drive off all volatile materials; error in weighing*

3. What errors in the experiment would cause the mineral percentage to be less than it should?

*Spillage, splattering of the ash, error in weighing*

EXPERIMENT 7

Chapters 1, 19 Exc. 12

# Fermentation and Apple Cider

Yeast cells secrete an enzyme that reacts with sugar to produce carbon dioxide. The reaction also produces either an acid, such as acetic acid, or an alcohol. Whether an acid or an alcohol is produced depends upon the type of sugar used and the kind of yeast present. In this experiment, you are interested in producing ethyl alcohol. This process is called *fermentation.* It is the basis for the alcoholic beverage industry.

You will ferment the carbohydrate found in apples to produce apple cider. Usually it is not necessary to add yeast to produce the reaction. Yeast is a natural contaminant of the atmosphere. Since the yeast is normally on the skin of the apple, you will not peel the apple, thus insuring that yeast is available for fermentation.

● Fermentation is the oldest known chemical process used by humans—it can be traced back at least 5000 years! The same process is used to make wine and in brewing. Fermentation usually stops when the alcohol content reaches 12%. At this concentration the growth of the yeast is inhibited. More concentrated solutions, such as those found in whisky and brandy (40-45%), are made by distillation of the fermented material.

● Yeast are capable of forming other products during the fermentation process, including hydrogen.

## PROBLEM

To prepare apple cider by fermentation.

● The reaction to produce ethyl alcohol is:

$$\underset{\text{glucose}}{C_6H_{12}O_6} \xrightarrow{\text{(enzymes)}} \underset{\text{ethyl alcohol}}{2C_2H_5OH} + \underset{\text{carbon dioxide}}{2CO_2}$$

### Materials

Ripe apples
Knife
Beaker
Mortar and pestle
Muslin cloth (or old clean pillow case)
125-mL Erlenmeyer flask with one-hole stopper to fit
Small length of glass tubing
Rubber tubing

## PROCEDURE

1. Cut the apples, with the skin left on, into small pieces. Crush the pieces using the mortar and pestle.
2. Put the crushed apples into a piece of muslin and squeeze as much juice as you can into the beaker. Keep twisting the ball of pulp tighter to get more juice.
3. Pour the apple juice into the Erlenmeyer flask, filling it to within 4 cm from the top.
4. Wet the glass tubing and gently insert it into the rubber stopper. Attach the rubber tubing to the glass tubing. Put the stopper in the flask.
5. Rinse out the beaker and fill it with water. Put the free end of the rubber tubing into the water.
6. Let the apple juice stand for 2 or 3 days. Look at it from time to time to see if there is any evidence of a reaction taking place. Check your "cider mill" after 2 or 3 days. Taste the cider. Remember, eating or tasting is almost never allowed in a science lab because many chemicals are poisonous, even in small amounts. On the few occasions when tasting is permitted, great care should be taken to be sure the equipment is clean. Equipment for making cider must be kept separate from other lab equipment. Be sure to wash it carefully before and after you use it.

**CAUTION** Be sure to use care in inserting the glass tubing into the rubber stopper. Moisten the glass tubing first with water or glycerine.

**CAUTION** Be sure that the rubber tubing is not clogged or stopped up. If it is, pressure will build up in the flask and it may explode. The stopper should be firmly inserted into the flask to prevent $CO_2$ from escaping around the stopper.

## QUESTIONS

**1.** Describe the cider produced in this fermentation reaction.

*The cider should be clear with a light brown color.*

**2.** What evidence is there that a gas was produced?

*The formation of gas is evident because of the bubbles forming in the beaker.*

**3.** How could you prove that $CO_2$ was produced?

*The gas can be tested with limewater. $CO_2$ gives a milky precipitate with limewater.*

**4.** Find out how the fermentation process is used to make alcoholic beverages.

*Students can find this in the library. They can also write any of the brewers for information. Most states have a Beverage Control Department and will be more than willing to provide additional material, including speakers, for your class.*

● Demonstrate the proper procedure for *tasting* substances in the laboratory—touch the finger to the cider and then to the tip of the tongue.

● As an optional extension to this experiment, you might like to have your students distill this alcohol solution. Consult a chemistry laboratory manual for instructions for distillation.

● This experiment is designed to give students experience in chemical analysis and introduce them to some compounds and chemical formulas.

EXPERIMENT 8

Chapters 4, 5

# What's In Baking Powder?

Baking powder is a mixture of at least two chemical compounds. One is bicarbonate of soda ($NaHCO_3$). This compound reacts with acids to produce a gas, carbon dioxide ($CO_2$), which causes bread to rise when it is being baked. The other compound, or compounds, react with water to form acids. The commonest acid-forming compound in baking powder is a tartrate—usually tartaric acid or potassium acid tartrate. Other possible compounds include a phosphate, such as calcium phosphate, and sodium aluminum sulfate, which is commonly called alum. According to government standards, baking powder must release at least 12% of its mass in the form of available $CO_2$.

● Potassium acid tartrate is potassium bitartrate.

## PROBLEM

To determine that baking powder produces carbon dioxide gas. To determine if a sample of baking powder contains a tartrate, a phosphate, or a sulfate as an acid-forming substance.

### Materials

● Old baking powder will not give very good results.

● The concentrations of the barium chloride and ammonium molybdate solutions are not critical.

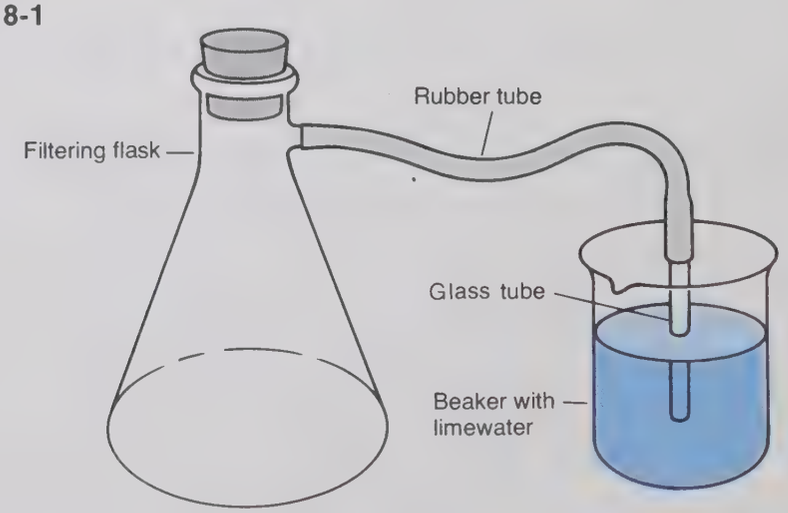
| | |
|---|---|
| Fresh baking powder | Test tubes |
| Vinegar | Barium chloride solution |
| Filtering flask | .1 *M* HCl |
| Rubber tubing | Ammonium molybdate solution |
| Glass tubing | $H_2SO_4$ (concentrated) |
| Beakers | Bunsen burner |
| Limewater | .1 *M* $HNO_3$ |
| Filter paper and funnel | Ringstand |

## PROCEDURE

### Part A: Preparing $CO_2$

**1.** Set up the apparatus as shown in Figure 8-1.

**Figure 8-1**



2. Place 2 g of baking powder in the flask. Connect the delivery tube so that it leads into the beaker containing limewater.
3. Mix 25 mL of vinegar and 25 mL of water in a small beaker. Add the entire 50 mL to the flask and quickly insert the stopper in the flask.
4. What happens to the limewater?

*It becomes milky.*

5. How do you know that $CO_2$ is responsible for this change? Slowly blow your breath through a plastic straw into another beaker containing limewater. Your breath contains a high concentration of $CO_2$. Compare the results with step 4.

*Results should be similar.*

**Part B: Determining the Acid-Forming Substance.**

1. Dissolve 10.0 g of baking powder in 50 mL of water.
2. After foaming has ceased, filter the solution through filter paper.
3. Use the filtrate for the tests.
4. Place 5 mL of the filtrate in a test tube.
5. Add 3 to 5 drops of barium chloride solution.
6. Check to see if a white precipitate forms. This is likely to be a sulfate.
7. If a white precipitate does form, add 3 to 5 drops of .1 *M* HCl. If the white precipitate is a sulfate, it will *not* dissolve. The only sulfate in baking powder would be sodium aluminum sulfate (alum).
8. In another test tube place 5 mL of the filtrate.
9. Add 3 to 5 drops of 0.1 *M* $HNO_3$.
10. Warm the solution in a hot water bath.
11. Add, drop-by-drop, a small amount of ammonium molybdate solution.
12. A phosphate will produce a bright yellow precipitate.
13. Place 5 mL of the filtrate in a clean evaporating dish.
14. Add 6 to 8 drops of concentrated $H_2SO_4$. **(CAREFUL!)**
15. Evaporate the solution completely over a small flame.
16. Check for a charred residue with an odor of burnt sugar. This indicates the presence of tartrate.

● Instruct students in the proper handling of acids.

## QUESTIONS

1. Based upon your results of this experiment, state the composition of your sample of baking powder.

*$CO_2$ plus some acid-forming compound which will vary.*

2. Is baking powder a *compound* or a *mixture*?

*A mixture*

3. Write out the equation leading to the formation of $CO_2$ in baking powder.

*Acid* + $NaHCO_3 \rightarrow CO_2 + H_2O$

4. Some gases can be collected by displacing a volume of water. It is difficult to collect $CO_2$ by this method. Why?

*It is extremely soluble in water. One liter of water at 20° C will dissolve .9 liter of* $CO_2$ *to form* $H_2CO_3$, *carbonic acid.)*

EXPERIMENT 9

Chapters 4, 5

# Decomposing Sodium Bicarbonate

Sodium bicarbonate (baking soda) is a substance commonly used in baking to cause bread to rise. This happens because sodium bicarbonate liberates a gas when it decomposes. If you performed Experiment 8, you saw that this gas is carbon dioxide, $CO_2$. But what else is produced?

In this experiment, you will decompose a solution of $NaHCO_3$ and a solid sample of $NaHCO_3$. You will then devise a scheme to determine what products are formed by the reaction. You should ask your teacher for whatever materials you need to assist you in your scheme to determine the products.

## PROBLEM

To investigate the decomposition of sodium bicarbonate and to determine the products.

### Materials

Sodium bicarbonate solution
Sodium bicarbonate (solid)
Litmus paper
250-mL beaker
Glass funnels
Test tubes
Small glass plate

● You should also have available for the students, upon request, the following:
Solid samples of NaOH, $Na_2O$ and $Na_2CO_3$
Miscellaneous items (splints, etc.)

## PROCEDURE

### Part A: Decomposing $NaHCO_3$ Solution

1. Prepare a solution of sodium bicarbonate by dissolving 5 g in 100 mL of water in a small beaker.
2. Test the sodium bicarbonate solution with litmus paper. Record your results in Table 9-1.
3. Bring the solution to a boil over a Bunsen burner. Then, place an inverted funnel over the top of the beaker so that the steam produced comes out like smoke from a stack.
4. Hold a cold glass plate above the inverted funnel and collect a few drops of the liquid as it condenses.
5. Test the liquid with litmus. Record your observations in Table 9-1.

**TABLE 9-1 • Testing $NaHCO_3$ Solution**

| TEST | OBSERVATIONS |
|---|---|
| $NaHCO_3$ Solution + Litmus | *Red litmus turns blue*<br>*Blue remains blue* |
| Condensed Liquid + Litmus | *Red remains red* |
| Remaining Solution + Litmus | *Red turns blue*<br>*Blue remains blue* |

6. Remove the funnel and continue to boil the solution until its volume is reduced by about ½.
7. Test the remaining solution with litmus paper. Record your observations.

### Part B: Decomposing the Solid Sodium Bicarbonate

1. Determine the mass of an empty test tube. Record the mass in Table 9-2.
2. Add enough solid sodium bicarbonate to cover the bottom of the tube.
3. Determine the mass of the tube plus sodium bicarbonate. Record the mass.
4. Heat the tube containing the sample. What happens?

_____

_____

5. Heat the tube strongly for about 3 minutes. Let the tube cool. Describe the contents of the tube.

_____

_____

6. When the tube is cool, determine the mass (with the product still in the tube). Record this mass.

● Encourage students to develop their own scheme to determine the composition of the product when sodium bicarbonate decomposes.

**TABLE 9-2 • Decomposing Solid $NaHCO_3$**

| | |
|---|---|
| Mass of Test Tube | |
| Mass of Test Tube + $NaHCO_3$ | |
| Mass of $NaHCO_3$ | |
| Mass of Test Tube + Residue | |
| Mass of Residue | |

### Part C: Determining the Product

1. There are several things that could possibly happen when sodium bicarbonate is heated. They are expressed by the following three reactions:

   **a.** $NaHCO_3\ (s) \rightleftharpoons Na_2O\ (s) + H_2O + CO_2$
   **b.** $NaHCO_3\ (s) \rightleftharpoons Na_2CO_3\ (s) + H_2O + CO_2$
   **c.** $NaHCO_3\ (s) \rightleftharpoons NaOH\ (s) + CO_2$

   Thus, the product could be either $Na_2O$, $Na_2CO_3$, or NaOH. Balance the three equations.

● Coefficients are:
a. 2, 1, 1, 2
b. 2, 1, 1, 1
c. 1, 1, 1

2. Design and carry out an experiment to determine which is the correct product. Ask your teacher for whatever material you need for your scheme. For your information, the formula masses of the compounds are:

   $NaHCO_3 = 84$ $\quad$ $NaOH = 40$
   $Na_2CO_3 = 106$ $\quad$ $Na_2O = 62$

## QUESTIONS

**1.** How does litmus react with the original solution? How does litmus react with the condensate? How does litmus react with the product remaining in the beaker? What can you conclude from this?

*See Table 9-1. Sodium bicarbonate ionizes to produce a basic solution.*

**2.** What substance is given off when solid $NaHCO_3$ is heated? How can you determine this?

*$CO_2$ test with limewater.*

**3.** Describe the quantitative test you designed and performed to predict the formula of the product remaining in the tube after $NaHCO_3$ is heated. Show your calculations.

Hopefully students will develop the following quantitative scheme:

1. What is the *ratio* of mass of the $NaHCO_3$ in this reaction? Obtain the weight ratio of $NaHCO_3$ to the residue as follows:

$$\frac{\text{Wt of residue}}{\text{Wt of original sample}} = \text{(ratio)} \frac{\text{residue}}{\text{sample}}$$

2. What is the ratio of residue to sample for the three possible products?

a. $\frac{Na_2CO_3}{NaHCO_3} = \frac{106}{168} = 0.63$

b. $\frac{NaOH}{NaHCO_3} = \frac{40}{168} = 0.24$

c. $\frac{Na_2O}{NaHCO_3} = \frac{62}{168} = 0.37$

3. Which ratio in (2) best fits the data from (1)?
Students ratio from (1) should be close to 0.63 for the correct product $Na_2CO_3$. This is a good opportunity to stress the Law of Conservation of Mass and the Law of Definite Composition!

● Borax is a very important mineral ore of boron. It is found in volcanic areas. About 85% of the world's supply of this mineral is found in the Death Valley area of California, apparently as deposits of past volcanic eruptions.



# Borax in the Doggie Biscuit

Many high carbohydrate foods such as crackers, ice-cream cones, cookies, wafers—and even dog biscuits contain either boric acid or borax as a preservative. Borax is tetraborate decahydrate. It has the formula $Na_2B_4O_7 \cdot 10H_2O$.

Read the label on boxes containing these types of foods. If the label indicates the presence of either boric acid or borax, there are simple tests that you can perform to identify these chemicals.

## PROBLEM

To test for the presence of boric acid or borax in food.

### Materials

Crackers, wafers, or dog biscuits containing boric acid or borax
Dilute HCl
Limewater
Turmeric paper
Litmus paper
Glass stirring rod
Mortar and pestle
Test tube
Crucible
Bunsen burner
Ringstand
Clay triangle

## PROCEDURE

1. Grind a small amount of the food sample to a fine powder in the mortar.
2. Place a 2 to 3 g sample into a crucible.
3. Using a dropper, add about 2 mL of limewater to the sample in the crucible until the sample is moist.
4. Test with litmus paper. The mixture should be basic.
5. Heat the sample until only a small amount of ash is left.
6. After the crucible cools, add 0.5 mL of HCl and stir with the glass rod until the solid goes into solution.
7. Moisten a piece of turmeric paper with the solution. Allow the test paper to dry by wrapping it around a test tube of hot water. The presence of boric acid or borax is indicated if the turmeric paper turns pink or red.

## QUESTIONS

1. Find out what other uses borax has. (Use your school library!)

*Borax is used in making pyrex glass, in certain glazes, to prepare boron, and so on.*

2. Find out the source of borax.

*See the teacher's note.*

EXPERIMENT 11

Chapter 5

# Laundry Bleach and Chlorine Gas

● This is a good introduction to ionic reactions. Stress the reversibility of the reaction by adding $H^+$.

Liquid household bleaches are usually a 5% solution of NaOCl, sodium hypochlorite. Usually these bleaching solutions are prepared by bubbling chlorine gas through a solution of sodium hydroxide.

$$Cl_2(g) + 2OH^- \rightleftharpoons ClO^- + Cl^- + H_2O$$

Notice that the hypochlorite ion, $ClO^-$, contains chlorine. In this experiment you will generate chlorine gas from bleach. You will do this by reversing the above reaction by adding HCl to the hypochlorite solution.

$$ClO^- + Cl^- + 2H^+ \rightleftharpoons Cl_2(g) + H_2O$$

● This is an extremely easy and safe way to produce $Cl_2$ gas. It is also a good way to produce chlorine water.

You can identify $Cl_2$ by its ability to bleach colors and dyes.

## PROBLEM

To prepare chlorine gas by decomposing laundry bleach. To prepare chlorine water and demonstrate its bleaching properties.

### Materials

Laundry bleach (sodium hypochlorite)
HCl
Samples of colored cotton cloth
Filtering flask and stopper
Test tube
Rubber tubing
Beaker

## PROCEDURE

### Part A: Preparing Chlorine Gas

1. Add 25 to 30 mL of laundry bleach to the filtering flask. Attach the rubber tubing to the side arm.
2. Working under a hood, place the open end of the rubber tube into the mouth of an inverted test tube.
3. Have your partner add 5 mL of HCl to the flask, stopper it quickly, and gently swirl the flask. Keep the end of the rubber tube inside the test tube.

**CAUTION** This should be done under a hood.

4. Is there evidence that $Cl_2$ is produced? Note any color and odor. Record your observations.

*$Cl_2$ will have a distinct light green color and an odor.*

**Part B: Preparing Chlorine Water**

1. Add 25 mL of laundry bleach to a 250–mL beaker.
2. Add 125 mL of distilled water to the bleach.
3. Place a piece of colored cloth in the diluted bleach solution. Record your observations.

*No reaction.*

4. Carefully add 10 mL of HCl to the dilute bleach containing the cloth.
5. Record your observations.

*The chlorine water will bleach the cloth.*

## QUESTIONS

1. Why should the preparation of chlorine gas be carried out under a hood?

*The gas is poisonous.*

2. Find out some other uses for chlorine.

*Bleach textiles and paper pulp, bacteriacide in drinking water*

*and swimming pools, preparation of drugs*

## EXPERIMENT 12

Chapter 5

# Preparing Oxygen Gas

● This experiment introduces the student to the use of a catalyst.

Laundry bleaches contain oxygen in the form of the hypochlorite ion, $ClO^-$. In this experiment you will decompose this ion and produce oxygen gas, $O_2$.

$$2ClO^- \text{—catalyst—}\longrightarrow O_2 + 2Cl^-$$

In order to form oxygen quickly, you will use a catalyst. A catalyst is a substance that speeds up the rate of a chemical reaction. The cobalt (II) ion, from a solution of cobalt (II) chloride, will act as the catalyst. You will test for the presence of oxygen by its ability to support combustion.

## PROBLEM

To produce oxygen gas from laundry bleach by using a catalyst.

● The following reaction can be effectively used with Chapter 15 (*Kinetics*) to show the effect of temperature and concentration on reaction rates. Add 3 mL of 0.2 *M* cobalt nitrate to 15 mL of bleach. At room temperature the bleach will decompose very quickly. By varying the temperature, the reaction rate will vary. You can easily estimate the volume of oxygen gas produced per 15 mL of various brands of bleach (a little consumer chemistry!)

### Materials

Laundry bleach (sodium hypochlorite)
Cobalt (II) chloride (solid)
Calcium hypochlorite solution
Filtering flask
Rubber tubing
Test tube
Shallow dish
Bunsen burner
Wooden splint

## PROCEDURE

1. Add water to the shallow dish until it is almost completely filled.
2. Attach a length of rubber tubing to the side arm of the filtering flask.
3. Fill a large test tube completely with water. With your thumb over the opening, invert the test tube into the dish of water. Be sure the test tube is still filled with water. Slip the end of the rubber tubing under the mouth of the test tube.
4. Have your partner add 100 mL of laundry bleach to the filtering flask.
5. Have your partner add 5 g of cobalt(II) chloride to the flask and stopper it quickly. Collect the gas produced by the reaction.
6. With your thumb over the opening, remove the test tube from the dish. While you are doing this your partner should ignite a wooden splint. Blow out the flame and insert the glowing splint (not burning) into the test tube.
7. Repeat the procedure using calcium hypochlorite instead of laundry bleach.

## QUESTIONS

1. What happens when cobalt(II) chloride is added to the flask?

*A black precipitate forms and oxygen gas bubbles from the precipitate.*

*(The precipitate is probably $Co_2O_3$, or a similar unstable oxide. It could decompose to form oxygen, then recombine with the hypochlorite ion.)*

**2.** Describe the test for oxygen.

*Oxygen supports combustion. The glowing splint will burst into flames.*

**3.** What is a catalyst?

*A substance that increases the speed of a chemical reaction but itself remains unchanged.*

EXPERIMENT 13 Chapter 5

# Smelling Salts, Fertilizer, and Ammonia

You can recognize ammonia by its unique odor. Ammonia is a very important ingredient in compounds such as fertilizers and smelling salts. In this experiment you will learn to identify ammonia, and then determine its presence in fertilizers, smelling salts, and protein.

## PROBLEM

To detect ammonia in fertilizer, smelling salts, and protein.

### Materials

Household ammonia
Beaker
Red litmus paper
1 *M* $NH_4Cl$
NaOH solution (concentrated)
Garden fertilizer (any commercial grade)
Ammonium carbonate (solid)
Smelling salts
Calcium oxide (solid)
Egg albumin (or other protein)
Bunsen burner
Test tubes

● Use the ammonia *without soap.*

● 1 *M* $NH_4Cl$—53.5 g/L

## PROCEDURE

### Part A: Identifying Ammonia and the Ammonium Ion

1. Place a small amount of household ammonia in a beaker. Fan a breeze over the beaker to carry some of the vapors to your nose as shown in Figure 13-1. Notice the odor of ammonia. (Caution—ammonia is toxic. Do not breathe in the vapors.) Hold a moist piece of red litmus paper over the

**Figure 13-1**

beaker. Record your observations. The odor and litmus tests will be used to identify ammonia.

*The odor is quite recognizable. The red litmus begins to turn bluish.*

2. Many ammonium salts react with strong bases to produce ammonia. This can be represented by the following net ionic equation.

$$NH_4^+ + OH^- \rightleftharpoons NH_3 + H_2O$$

Thus, you can identify $NH_4^+$ by the same odor and litmus reaction as for $NH_3$ when the $NH_4^+$ reacts with a strong base. Place 5 mL of 1 *M* $NH_4Cl$ in a test tube. Perform the odor and litmus tests. Record your observations.

*There should be no detection of ammonia.*

3. Add 1 mL of concentrated NaOH solution to the tube. Again perform the odor and litmus tests. (Do not allow the litmus to touch the test tube.) If no ammonia is detected, gently warm the test tube over a low Bunsen burner flame. Record your results.

*The odor of ammonia should be detected; red litmus turns blue.*

● Some ammonium compounds, such as ammonium chloride, are stable. However, some, like ammonium carbonate, decompose very readily to ammonia and carbon dioxide. $(NH_4)_2CO_3 \rightarrow 2NH_3 + CO_2 + H_2O$

**Part B: Testing Materials for the Presence of Ammonia**

1. Place about .1 g of garden fertilizer in a test tube. Add 1 mL of concentrated NaOH. Perform the odor and litmus tests. Record your results.

*The odor of ammonia should be detected; red litmus turns blue.*

2. Smelling salts often contain ammonium carbonate, $(NH_4)_2CO_3$ as an active ingredient. Perform the odor and litmus tests. Record your results.

*The odor of ammonia should be detected; red litmus turns blue.*

3. Nitrogen in protein is changed to $NH_3$ when heated with a strong base. Place 1 mL of concentrated NaOH in a test tube. Add a small amount of calcium oxide, CaO (about the size of a pea). Gently mix the two until the CaO dissolves. Add 1 mL of egg albumin, or another protein sample. Check the tube for the odor of ammonia. Test with moist red litmus for the presence of ammonia. Record your results.

*The odor of ammonia should be detected; red litmus turns blue.*

## QUESTIONS

1. Indicate which of the substances tested contained or produced ammonia.

*All of them*

2. Using this same technique, describe tests you could perform on other substances containing, or suspected of containing, ammonia or ammonium ion.

*Student answers will vary. Encourage innovation!!*

## EXPERIMENT 14

Chapter 5

# Baking Soda, Chalk, and $CO_2$

● This is a good experiment to show students how a compound, such as $CO_2$, can be produced from many sources. It also introduces the student to solubility and the formation of precipitates in solution.

Carbon dioxide, $CO_2$, is a colorless, odorless gas that is produced by compounds containing the carbonate group, $CO_3^{-2}$. It is also exhaled by humans and animals. Compounds containing $CO_3^{-2}$ react with acids to release $CO_2$, as in the following equation involving sodium bicarbonate.

● You might like to do this experiment in conjunction with the study of acids and bases.

$$NaHCO_3 + H^+ \rightleftharpoons CO_2 + H_2O + Na^+$$

You can easily test for the presence of $CO_2$, since it forms insoluble compounds when it reacts with substances such as barium hydroxide, $Ba(OH)_2$, in solution.

$$CO_2 + Ba(OH)_2 \rightleftharpoons BaCO_3(s) + H_2O$$

Barium carbonate is a white precipitate.

## PROBLEM

To test for the presence of $CO_2$ when substances containing the carbonate or bicarbonate group react with acids.

### Materials

| | |
|---|---|
| Sodium bicarbonate | 2 *M* HCl |
| 18 *M* $H_2SO_4$ | Lemon juice or sour milk |
| 1 *M* $Ba(OH)_2$ | Glass slide or small glass plate |
| Vinegar | Test tubes |

● Concentrated $H_2SO_4$ is 18 *M*.

● A 1 *M* solution of $Ba(OH)_2$ is needed—171.3 g/L.

## PROCEDURE

### Part A: Testing Baking Soda for $CO_2$

1. Place .1 g of fresh baking soda in a small, dry test tube.
2. While you are doing this, your lab partner should place a drop of $Ba(OH)_2$ on a clean, dry microscope slide.
3. Now, add 3 to 5 drops of $H_2SO_4$ to the test tube. Immediately turn the slide over so that the drop of $Ba(OH)_2$ is hanging upside down. Place the slide over the mouth of the test tube so that it is completely covered. The drop should be centered over the mouth of the test tube. Record your observations.

*A white precipitate forms in the drop.*

4. Repeat the experiment using vinegar instead of $H_2SO_4$. Repeat the experiment using sour milk or lemon juice instead of $H_2SO_4$. Record your observations.

*A white precipitate should form in both cases.*

● You might also like to have students test other common substances such as egg shell, limestone, oyster shells, and so on.

**Part B: Testing Chalk for $CO_2$**

1. Place a small piece of chalk, about the size of a pea, in a test tube.
2. Add 3 to 5 drops of HCl.
3. Test for the presence of $CO_2$ with the hanging drop of $Ba(OH)_2$ as you did before. Record your observations.

*A white precipitate forms.*

## QUESTIONS

1. What is the white precipitate that causes the drop of $Ba(OH)_2$ to become cloudy?

*Barium carbonate, $BaCO_3$*

2. When doing this experiment, you must be careful not to breathe on the drop of $Ba(OH)_2$. Why?

*$CO_2$ in the exhaled breath will cause a precipitate to form.*

## EXPERIMENT 15

Chapters 5, 15

# Sulfur Dioxide in Foods

● This experiment gives the student experience with precipitation, solubility, and calculating percentage composition of a formula.

Sulfur dioxide, $SO_2$, is a serious air pollutant and a noxious gas when present in high concentrations. However, it is useful in preserving foods—especially fruits. Sulfur dioxide is also used to bleach certain foods. Sugar, for example, is white because of the bleaching properties of $SO_2$. Sulfur dioxide is especially useful in preserving dried fruit, such as apples. It prevents the growth of bacteria and mold. If it were not for the $SO_2$ used to preserve raisins, they would arrive at the supermarket in one moldy lump! Sulfur dioxide also acts as an antioxidant and is used to prevent the brown color one sees with cut and bruised fruit. Apples, apricots, peaches, pears, and figs are dried and preserved with $SO_2$. Although most of the $SO_2$ is removed, there is some residual $SO_2$ on and in the fruit. You will test for the presence of this residual $SO_2$. You will isolate $SO_2$ by precipitating it in the form of the insoluble compound barium sulfate, $BaSO_4$.

● If you wish to discuss the role of $SO_2$ as an air pollutant, you can refer to any college level chemistry text. Students should not conclude that $SO_2$ is harmful when used as a preservative, but should be impressed with their ability to determine its presence and its concentration.

## PROBLEM

To determine the amount of $SO_2$ (as S) present in dried fruits.

● Students may need to review redox reactions and net ionic reactions before doing this experiment.

### Materials

Dried fruit (apples, apricots, and so on)
(Check to see if the package indicates that $SO_2$ was used as a preservative)
3 *M* $BaCl_2$
3% $H_2O_2$
Filter paper (fine grade)
Funnels
Cheesecloth
500-mL beakers
Spatula or stirring rod

● $BaCl_2$ solution—62.47 g of $BaCl_2$ in 100 mL of solution. Place this in several small dropper bottles.

## PROCEDURE

1. Determine the mass of several slices of dried apples, or 3 small apricots. Record the mass in Table 15-1.
2. Place the dried fruit in a 500-mL beaker and just cover it with distilled water.
3. Allow the fruit to remain in the water for 15 minutes. As the fruit absorbs water, it will swell. Add more water to restore the original volume. (Best results will be obtained if the fruit is allowed to soak overnight.)
4. Place a large piece of cheesecloth (or several unfolded gauze squares) in a large funnel. Place a beaker under the funnel. Pour the liquid and the fruit into the cheesecloth. Using a spatula or stirring rod, squeeze the fruit to extract as much juice as possible.
5. Add 50 mL of 3% $H_2O_2$ to the filtrate. This will oxidize the $SO_3^{-2}$ to sulfate ion, $SO_4^{-2}$.

$$SO_3^{-2}\,(aq) + H_2O_2 \rightleftharpoons SO_4^{-2}\,(aq) + H_2O$$

● Review solubility rules.

6. It is now necessary to separate the sulfate ion from solution. This is done by precipitating it as barium sulfate, $BaSO_4$. Add $BaCl_2$ solution drop-by-drop. Notice the formation of a white precipitate. Continue adding barium chloride until precipitation stops. We will assume that all of the $SO_4^{-2}$ has precipitated.
7. Prepare to filter the precipitate of barium sulfate. Select a filter paper that has fine pores. If the mass per sheet is not given on the filter paper box, determine the mass of the sheet and record it in Table 15-1.
8. Fold the filter paper and place it in a funnel. Place a beaker under the funnel.
9. Pour the liquid and precipitate onto the filter paper. Be sure to get all of the barium sulfate on the paper. If any precipitate remains in the beaker, add water to the beaker, swirl, and pour it onto the filter paper.
10. Carefully remove the filter paper containing the filtrate and set it aside to dry.
11. After the paper is dry, determine the mass of the paper and filtrate. Record the information and complete Table 15-1.

**TABLE 15-1 • Sulfur Content of Fruit**

| | |
|---|---|
| Mass of Fruit | |
| Mass of Filter Paper | |
| Mass of Filter Paper & Precipitate | |
| Mass of Precipitate | |

## QUESTIONS

1. The precipitate, $BaSO_4$, has a molecular mass of 233. Sulfur has an atomic mass of 32. Determine the percent of sulfur in $BaSO_4$.

*32/233 = .137 = 13.7%*

2. Determine the mass of sulfur in your precipitate.

*Mass of precipitate x .137*

3. Determine the percent of sulfur in the fruit.

*Mass of sulfur/Mass of fruit*

4. How could you determine the mass of $SO_2$ in the fruit?

*Since sulfur constitutes 50% of the mass of $SO_2$, the mass of $SO_2$ in the fruit should be twice the mass of sulfur present.*

# What Makes Popcorn "Pop"?

If you examine a single kernel of popcorn, and compare it with a grain of ordinary corn, two differences can be easily noted. First, the kernel of popcorn is more rounded than that of ordinary corn. Second, the coat on the popcorn is thicker and tougher. Both of these features are important in the "popping" of popcorn.

Even though popcorn seems to be dry, each kernel contains a small amount of water. When the popcorn is heated, the water forms steam. When sufficient pressure is built up within the kernel, the thick coat bursts, allowing the steam to escape. It is the heat and sudden expansion that causes the white material in the seed to "fluff up." The thick seed coat holds the steam in until just the right conditions occur for popping.

## PROBLEM

To investigate the "popping" of popcorn.

● Allow the students to discover that the gas laws apply to the popping of popcorn. The increase in temperature causes a tremendous increase in pressure in a small, confined area. The tough coat allows the pressure to build up enough to pop the shell. It also allows enough time for the white, pulpy material to get hot enough to fluff-up. The popcorn would not "pop" if there wasn't a bit of water inside the kernel. This forms steam (gas) that expands and builds up pressure in a constant volume.

### Materials

| | |
|---|---|
| Popcorn | Flask and stopper |
| Ordinary corn | Bunsen burner |
| Cooking oil | Pin or other sharp pointed object |
| Beaker | Hotplate |

## PROCEDURE

1. Select samples of about 10 to 15 kernels of popcorn for each of three tests. Determine the mass of each sample of kernels. Record the mass in Table 16-1. Determine the average mass of a single kernel.
2. Cover the bottom of a flask with cooking oil. Pour out any excess. Place one sample of popcorn kernels in the flask. Place a loose fitting stopper in the flask. Gently heat the flask with a Bunsen burner. Shake the popcorn occasionally. Continue until the popcorn "pops."
3. Determine the mass of the popped popcorn. Record the mass. Determine the average mass of a single popped kernel.

● Emphasis can be placed on recording data and averaging to obtain the mass of ■ single kernel of corn. Be sure to point out the limitations in such procedures.

**TABLE 16-1 • Popping Popcorn**

| | TEST 1 | TEST 2 | TEST 3 |
|---|---|---|---|
| Mass of Unpopped Sample | | | |
| Mass of Single Kernel | | | |
| Mass of Popped Sample | | | |
| Mass of Popped Kernel | | | |

● If the kernel is pierced or cracked, the pressure will not build up, and the corn will not pop.

4. Treat the second sample as you did the first. However, before heating the kernels in this sample, carefully pierce each sample with the needle or pin. Be sure to completely pierce the seed coat, but do not crack it. Record your data.
5. Place the third sample of popcorn kernels in a beaker. Do not put any cooking oil in the beaker. Place the beaker in an oven or on a hot plate set at a low temperature. Gradually increase the heat over a 30 to 40 minute period. Record your results.

## QUESTIONS

1. How does the average mass of a single kernel of corn compare before and after popping?

2. What happens to the popcorn that has been pierced prior to popping?

*It probably will not pop.*

3. What happens to the popcorn that is heated slowly rather than quickly?

*It should make no difference.*

4. What effect does the cooking oil have in the first sample?

*It evenly distributes the heat to the popcorn kernels.*

## EXPERIMENT 17

Chapter 5

# The Molecular Mass of Butane

● Use this experiment to introduce the gas laws.

The Avogadro hypothesis tells us that the molar volume of any gas, at 273 K and 760 Torr, is 22.4 liters. You can use this information to determine the molecular mass of any gas. In this experiment, you will determine the molecular mass of the gas *butane*, commonly found in inexpensive pocket lighters.

Butane has the formula $C_4H_{10}$. You will release butane from a pocket lighter and allow it to displace a measured volume of water. From the displaced volume, the measured temperature, and the pressure, you can easily calculate the molecular mass of butane gas.

● Since the difference in mass of the lighter before and after the experiment is so small, a large volume (200 mL) is required for accuracy.

## PROBLEM

To determine the molecular mass of butane.

### Materials

250-mL Erlenmeyer flask
Glass plate
Trough
Butane lighter
Thermometer
Marking pencil
Graduated cylinder

● Any of the disposable butane lighters will work. First, you should remove the flint and striker. Plug the small hole left by the flint with modeling clay (just a dab to cover the opening). Unless the entire class is doing the experiment at one time, one lighter will do. If you do not have graduated flasks, use ordinary Erlenmeyer flasks, but have wax pencils available for the students to mark the volume.

## PROCEDURE

### Part A: Collecting the Butane

1. Fill a 250-mL graduated Erlenmeyer flask completely full of water. Cover the top with a small glass plate.
2. Fill a trough or some other suitable container with water. Holding the glass plate in place with your finger, carefully turn the flask upside down and place it in the trough. Be sure that no air bubbles enter the flask. Remove the glass plate.
3. Your teacher will have some lighters for your experiment. Select a lighter and wipe it dry and clean with a paper towel. Check to see that it is turned up to give the highest gas flow.
4. Place the cleaned lighter on a balance pan and carefully determine the mass to the nearest hundredth of a gram. (Estimate as closely as possible.) Record the mass in Table 17-1.
5. Carefully hold the butane lighter under the water. Make sure that the gas opening is beneath the mouth of the inverted flask. Press the release lever, being sure that the gas bubbles enter the flask. (See Fig. 17-1 on page 36.)
6. Continue to hold the lever down until you have collected approximately 200 mL of gas.
7. Remove the butane lighter, dry it with a paper towel, and set it aside.

● It is most important that students record volume and mass as accurately as possible. When you are discussing the calculations, show them how much 1 mL or .1 g affects the answer. Be sure there are no air bubbles in the flask, and be sure that all of the gas bubbles enter the flask.

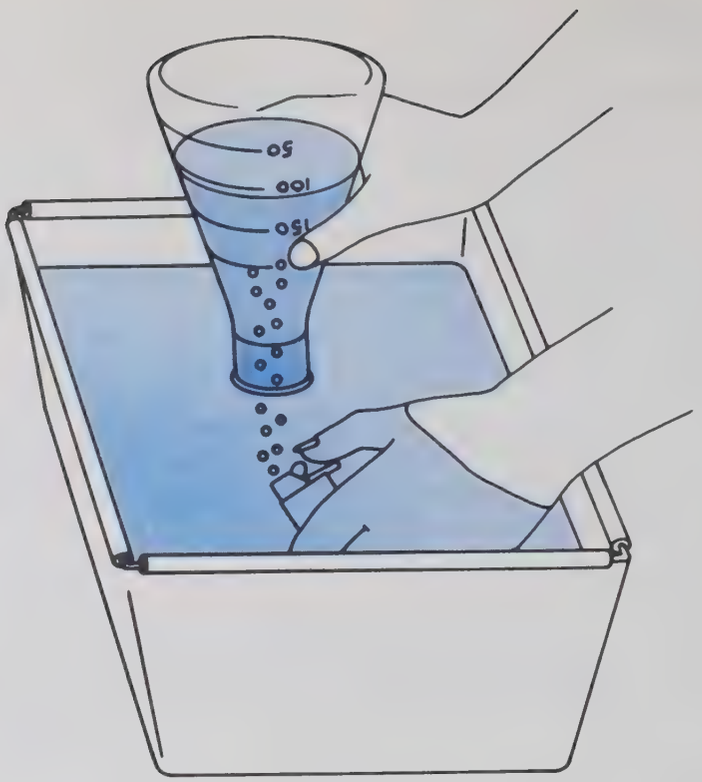



**Figure 17-1**

- Demonstrate the procedure for equalizing the water levels prior to the experiment.

8. Do not remove the gas-filled flask from the water. Carefully raise or lower the flask until the level of water *inside* the flask is the same as the level of water *outside* the flask. This is to equalize the pressure inside and outside the flask.

- Actually, if students can mark the gas level, they can use any glass container to collect the gas.

9. Read the volume of gas on the flask. Record the volume in Table 17-1. If you cannot read the volume accurately, mark the water level on the flask with a marking pencil. Remove the flask from the water and fill it to your recorded mark with water. Now, pour the measured volume of water into a graduated cylinder where you can read the volume more accurately. Record this volume in Table 17-1.

- One lab partner can weigh the lighter while the other determines the volume of gas in the flask.

10. Determine the mass of the butane lighter as you did before. Record the mass in Table 17-1.

- If the physics department doesn't have a barometer, call your weather bureau and ask for barometric pressure. If they give it to you in inches, refer to any handbook to convert it to mm of Hg (Torr).

11. Measure and record the temperature of the water in the trough. Record the barometric pressure—your teacher will give you this information.

**TABLE 17-1 • Butane Collection**

| | |
|---|---|
| Mass of Lighter before Experiment | *21.34 g* |
| Mass of Lighter after Experiment | *20.88 g* |
| Mass of Butane | *.46 g* |
| Volume of Gas Collected | *200 mL* |
| Temperature (°C) | *19* |
| Pressure (Torr) | *756* |

## Part B: Calculating Molecular Mass

1. Determine the partial pressure of the butane gas. The gas in the flask is not all from the butane lighter. There is also water vapor. Since you are only interested in the pressure due to the butane, you must *subtract* the pressure due to water vapor. Table 17-2 gives the partial pressure of water vapor at various temperatures. Select the vapor pressure that corresponds to the temperature which you recorded in Table 17-1. Subtract this from the pressure reading in the table.

$$P_{Butane} = P_{room} - P_{H_2O}$$

2. Determine the volume the butane would have at 760 Torr (standard pressure).

$$\text{Volume} = V_{measured} \times \frac{P_{butane}}{760 \text{ Torr}}$$

3. Determine the volume the butane would have at 273 K (standard temperature).

$$\text{Volume} = \text{Volume (from step 2)} \times \frac{273 \text{ K}}{273 \text{ K} + \text{Temp in table}}$$

4. You now have a certain volume of gas produced by a certain mass of gas. Now ask, "What mass of the gas would produce 22,400 mL?"

$$\frac{\text{Mass of gas (g)}}{\text{Volume of gas (mL)}} = \frac{\text{"X" of gas}}{\text{22,400 mL of gas}}$$

or,

$$\text{"X"} = \frac{\text{(Mass of gas) (22,400 mL gas)}}{\text{Volume of gas (mL)}}$$

5. How close did you come?

   **a.** Record your answer.

*56.6 g*

   **b.** Calculate the molecular weight of butane, $C_4H_{10}$.

*58.1 g*

● Sample calculations:

$P_{Butane} = 756 \text{ Torr} - 16.5 \text{ Torr} = 739.5 \text{ Torr}$

$V = 200 \text{ mL} \times \frac{739.5 \text{ Torr}}{760 \text{ Torr}}$

$V = 194.6 \text{ mL}$

$V = 194.6 \times \frac{273 \text{ K}}{273 + 19}$

$V = 181.94 \text{ mL}$

$\text{"X"} = \frac{(0.46 \text{ g}) \quad (22{,}400 \text{ mL})}{181.94 \text{ mL}}$

"X" = 56.6 g (per mole)

● If students are very careful, they can come within 3 to 5% of the molecular mass of butane. After performing the experiment and doing the calculations, they may wish to repeat the experiment. This experiment gives practice in calculations, molar volume, gas laws, and good lab technique!

**TABLE 17-2 • Vapor Pressure of Water at Various Temperatures**

| TEMPERATURE (°C) | PRESSURE (Torr) | TEMPERATURE (°C) | PRESSURE (Torr) |
|---|---|---|---|
| 15 | 12.8 | 23 | 21.0 |
| 16 | 13.6 | 24 | 22.4 |
| 17 | 14.5 | 25 | 23.7 |
| 18 | 15.5 | 26 | 25.2 |
| 19 | 16.5 | 27 | 26.7 |
| 20 | 17.5 | 28 | 28.3 |
| 21 | 18.6 | 29 | 30.0 |
| 22 | 19.8 | 30 | 31.8 |

# Why Does Silverware Tarnish?

Everyone who uses silverware or silverplate has noticed that it often tarnishes, or turns dark. This occurs when silverware has been exposed to the air for a long time. Contact with certain substances in food can also cause silverware to tarnish. The tarnish results when the silver reacts with sulfur, forming silver sulfide, $Ag_2S$. In this experiment you will expose silver to egg yolk. Egg yolk contains hydrogen sulfide, $H_2S$. Hydrogen sulfide is also present in small concentrations in the air. The hydrogen sulfide is oxidized to sulfur, which then reacts with silver to produce the tarnish. You can also cause a silver spoon or other article to tarnish by wrapping a rubber band around it. The sulfide in the vulcanized rubber has the same effect as the $H_2S$ in egg yolk.

● The reaction is $2Ag + S \rightleftharpoons Ag_2S$. This is a good time to introduce, or stress, oxidation-reduction reactions.

## PROBLEM

To investigate the process of tarnishing of silver.

### Materials

● Although you can tarnish the silver by using other sulfides, the use of egg yolks gives the experiment a more practical and everyday twist.

Silverware or silverplate (brought from home)
Egg
Beaker

## PROCEDURE

● Demonstrate the method of separating an egg from its yolk.

1. Separate the yolk from the egg white in a beaker.
2. Dip the silver article into the egg yolk and rub the yolk over the article.
3. Set the silver article aside for 10 to 15 minutes.
4. Wash the silver article gently.

## QUESTIONS

● Most students will suddenly realize one reason why their own silverware tarnishes—as a result of the food with which it comes in contact. Of course, most damage comes from $SO_2$ in the air. This gives you an opportunity to discuss problems of air pollution.

1. What evidence is there that the silver has tarnished?

*A black area appears where the silver was exposed to the egg yolk.*

2. Write the equation for the reaction between silver and sulfur to produce silver sulfide.

*$2Ag + S \rightleftharpoons Ag_2S$*

● This experiment is a good way to emphasize practical chemical reactions.

# Removing Tarnish From Silver

Silver is often cleaned by using a paste in which a mild abrasive is dispersed. This cleansing agent wears away the tarnish and exposes the silver beneath it. In this experiment you will clean silver *chemically*, rather than physically. You can use the tarnished silver from the previous experiment, or you can bring tarnished silver, silverplate, jewelry, and so on from home. In fact, you can easily do this experiment at home. The reaction for this experiment is:

$$2Al + 3Ag_2S + 6H_2O \rightleftharpoons 6Ag + 2Al(OH)_3 + 3H_2S$$

## PROBLEM

● Encourage students to do this experiment at home.

To chemically remove tarnish from silver.

### Materials

Sodium bicarbonate (baking soda)
Scrap aluminum (such as a disposable TV dinner pan)
Tarnished silver
Large metal pan
Tongs

● Pie pans and aluminum wrap will work well.

## PROCEDURE

1. Prepare a sodium bicarbonate solution by placing 7 to 8 tablespoons of sodium bicarbonate in 2 liters of water.
2. Place the solution in a large pan. Place the aluminum in the pan so that it remains on the bottom.
3. Heat the solution slowly, but do not boil it.
4. While the solution is hot, immerse the tarnished silver into the solution. Hold it in place with a pair of tongs. *The tarnished silver must touch the aluminum in the bottom of the pan.*
5. After the silver is cleaned, remove it from the sodium bicarbonate solution and wash it in running water.

● A heating source will be necessary.

● If the piece of silver to be cleaned is too large to fit in the pan, clean one side or section at a time—rotating it through the solution.

## QUESTIONS

● If the solution doesn't remove the tarnish, add more baking soda.

1. What evidence is there that this is a chemical reaction?

*Bubbles of gas form, the tarnish (which is a sulfide) is removed, and so on.*

2. What gas was evolved?

*$H_2S$, although you may not be able to detect it by its characteristic "rotten egg" odor.*

● This is a good experiment to use to illustrate differences in physical properties. By changing conditions, we are changing the physical state of ions and molecules. It is also a good time to begin a discussion of atomic structure and crystal formation.

**EXPERIMENT 20**

Chapters 6, 10, 16

# Making Crystals

Atoms, molecules, and ions can fit together in a regular pattern to form a *crystal*. There are several types of crystals and crystalline structures. In this experiment you will investigate crystals of sugar and crystals of salt. You will produce salt crystals by allowing ions of sodium and chlorine to come together in solution. You will produce sucrose crystals (called rock candy) by allowing sugar molecules to come together in solution.

## PURPOSE

To prepare and compare crystals of salt and sugar.

### Materials

Sugar
Salt
String
Lead weights
Beakers or glasses
Hotplate

● String used to wrap packages works best. Don't use thread or heavy cord. Fishing weights or any weights that will hold the string taut will work.

## PROCEDURE

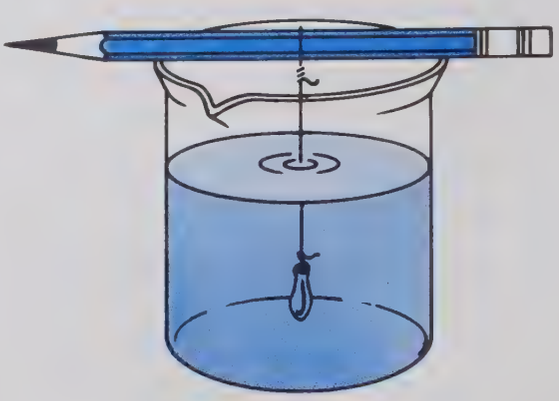
### Part A: Preparing Salt Crystals

1. Fill a 250–mL beaker ¾ full of water.
2. Add salt to the water until no more will dissolve.
3. Heat the solution, but do not boil it.
4. While the solution is hot, add salt until no more will dissolve. Remove the beaker from the hotplate.
5. Tie a weight on the end of a string.
6. Suspend the string in the beaker, as shown in Figure 20-1. Support the string in the middle of the beaker with a pencil, soda straw, or wood splint.
7. Put the beaker where it will not be disturbed for a few days. Make observations daily to see if crystal growth is occurring.

● Students can transfer the solution to a drinking glass, if desired.

● It is a good idea to set the beakers on an aluminum pie-plate, as they may become messy as the solids crystallize.

**Figure 20-1**



**Part B: Preparing Sugar Crystals**

1. Boil 150 mL of water in another 250–mL beaker.
2. Add as much sugar to the hot water as will dissolve.
3. Repeat steps 5 through 7 from Part A.
4. Compare both types of crystals.

## QUESTIONS

1. Describe the results of this experiment.

*Crystal should form in both beakers.*

2. What purpose does the string serve?

*The string acts as a "nucleus" for crystal formation.*

3. What are the characteristics of each type crystal?

*Answers will vary.*

● Both solutions are supersaturated, meaning that they contain more substance in solution than they can hold under ordinary conditions. When conditions return to normal, when the solution cools and evaporates, crystals form on the string.

## EXPERIMENT 21

# Boiling Water in a Paper Cup

Heat energy absorbed by matter is converted into molecular kinetic energy. This is indicated by a rise in the temperature. Different substances respond differently to the absorption of the same amount of heat energy. When water is placed in a paper cup (or a paper bag) and heated, the water absorbs heat much faster than the paper. Thus, water will continue to absorb the heat until it reaches its boiling temperature. It will then continue to absorb heat for the boiling process. In effect, water will pull all of the heat from the flame, leaving none available to burn the paper cup.

## PROBLEM

To boil water in a paper cup.

### Materials

- Dixie cups work well. The small, unwaxed paper cups used in the hospital to dispense pills work better. The paper cups must *not* be wax coated.

Paper cup
Bunsen burner
Ringstand
Clay triangle

## PROCEDURE

1. Place a paper cup inside a clay triangle or directly on a ringstand.
2. Fill the paper cup *completely to the brim* with water.
   - If the cup is not filled completely with water, the top edge will char.
3. Heat the paper cup with a Bunsen burner, fanning the flame over the entire cup.
   - Caution the students about using the burner and avoiding accidents with the boiling water.
4. Continue the process until the water boils. Shut off the burner and allow the water to cool.

## QUESTIONS

1. Examine the paper cup after the water cools. Record your observations.

*The cup should be undamaged.*

2. Explain your observations.

*The water absorbed most of the heat for the boiling process, leaving none to burn the cup.*

**EXPERIMENT 22** Chapters 6, 7 Exc. 11

# Analyzing Razor Blades for Chromium

The stainless steel generally used in razor blades contains iron combined with other metals, such as nickel and chromium. In this experiment you will recover chromium from a razor blade and determine the percent of chromium in the blade. The chromium will be converted to lead chromate, $PbCrO_4$, a bright yellow, insoluble substance.

● Announce this exercise several days in advance so students can ask their fathers to save their discarded blades.

● This experiment is intended to give the student experience with quantitative techniques—precipitating, filtering, determining mass, and calculating percent composition of an element in a compound. The experiment is easy, although it involves several steps; it includes many colored compounds, and the results are easily reproducible.

## PROBLEM

To estimate the percent of chromium in a razor blade.

● Chromium content of steel varies from 4 to 20%.

### Materials

Razor blade (used double-edged)
HCl (concentrated)
25% KOH
10% $H_2O_2$
Glacial acetic acid
.50 *M* lead acetate
250-mL beaker
400-mL beaker
Ice
Graduated cylinder
Filter paper
Bunsen burner
Ringstand
Funnel
Test tube
Stirring rod

● Lead acetate solution—133 g $Pb(C_2H_3O_2)_2$/liter.

## PROCEDURE

1. Use a double-edged razor blade for this experiment. Wrap a towel around the blade and *carefully* snap it in two.

**CAUTION** Show students how to break the blade.

2. Determine the mass of one of the halves to the nearest 0.01 g. Record the mass in Table 22-1.

● The ½ blade should weigh about 0.5 g.

3. Place the sample in a large test tube and add 15 mL of HCl. This should be done under a hood. Handle the acid carefully!

● HCl fumes will be given off.

4. Heat the tube very gently. What do you observe as the tube is heated?

*The solution is dark green due to the presence of chromium(III) ion.*

● $2Cr + 6HCl \rightarrow 3H_2 + 2CrCl_3$

5. Carefully pour the contents of the test tube into a clean 250-mL beaker. Rinse the tube with 10 mL of distilled water, adding the rinse water to the beaker. Nest the beaker in a bed of crushed ice.

● Demonstrate how to use a wash bottle to rinse a precipitate.

6. Slowly, while stirring, add 50 mL of 25% KOH to the beaker.
7. Slowly, with constant stirring, add 30 mL of 10% $H_2O_2$. What happens?

*The solution will bubble and froth.*

● $CrCl_3 + 4KOH \rightarrow KCrO_2 + 3KCl + 2H_2O$

$2KCrO_2 + 3H_2O_2 + 2KOH \rightarrow 2K_2CrO_4 + 4H_2O$

8. Continue to stir until the solution stops foaming.
9. Remove the beaker from the ice and continue to stir until no more bubbles are formed.
10. Heat the solution slowly until it begins to boil. Boil it for about two minutes. What is the appearance of the solution now?

*Yellow, due to the presence of chromate ion, $CrO_4^{-2}$*

● Heating will decompose any remaining peroxide.

11. Remove the beaker from the flame. Again while stirring, slowly add 15 mL of glacial acetic acid.
12. Allow the solution to cool. Filter it into a 400-mL beaker. Rinse the precipitate with three 5-mL portions of distilled water. Add the rinse to the filtrate. If the filtrate is still yellow, you must pour it back into the 250-mL beaker and mix it again with glacial acetic acid. After mixing, refilter it.

**CAUTION** Lead acetate is very toxic. At one time it was used as a yellow pigment in paints. It is now prohibited in certain items, especially children's toys and household products. Lead poisoning often results from ingesting lead compounds when drinking from porcelain containers containing them.

13. Add 20 mL of 0.5 *M* lead acetate to the filtrate in the large beaker. Heat the solution until it boils, then cool it using an ice bath. What happens?

*A precipitate forms.*

14. Determine the mass of a piece of filter paper and record the mass in Table 22-1. Filter the precipitate. The filtrate should be clear. If it is cloudy, filter it again through the same piece of filter paper. Wash the precipitate with 10 mL of water.

● If an oven is available, the paper can be dried at about 100°C.

15. Place the paper with the precipitate in an out-of-the-way place and allow it to dry over night.
16. Determine the mass of the dried paper containing the precipitate ($PbCrO_4$). Record your data in Table 22-1.

● $K_2CrO_4 + Pb(C_2H_3O_2)_2 \rightarrow PbCrO_4 \downarrow + 2KC_2H_3O_2$

**TABLE 22-1 • Chromium**

| | |
|---|---|
| Mass of ½ Blade | |
| Mass of Filter Paper | |
| Mass of Filter Paper + $PbCrO_4$ | |
| Mass of $PbCrO_4$ | |

## QUESTIONS

1. The formula mass of $PbCrO_4$ is 323.19. The atomic mass of Cr is 52.0. Determine the percent of Cr in $PbCrO_4$.

*16%*

2. Determine the mass of Cr in your precipitate.

*Mass of $PbCrO_4$ x .16 =*

3. What is the percent of Cr in the razor blade?

*Mass of Cr/mass of ½ razor blade.*

**EXPERIMENT 23** Chapters 6, 7 Exc. 11

# Determining the Iron Content in Foods

● This is an experiment in analytical technique. The experiment gives the students experience with colorimetric analysis and experience with the technique of diluting from a standard solution.

Iron is a very important component of the diet. It is used to make the pigment hemoglobin which is found in red blood cells. In this experiment you will develop a test for the presence of iron. You will also make a quantitative estimate of the amount of iron present in a sample of food. This experiment will involve a *colorimetric* determination. You will prepare a set of standard solutions with known concentrations of iron. The standards will vary in color from deep red to colorless. You can determine the amount of iron in your food sample by comparing the color of the prepared sample with the colors of the standards. This is a common analytical technique.

Iron, specifically $Fe^{+3}$, reacts with KSCN to form a deep red colored solution. The net ionic reaction is:

$$Fe^{+3} + SCN^{-} \rightleftharpoons FeSCN^{+2}$$
(colorless) (deep red)

The intensity of color produced varies with the concentration of iron in the sample.

● This should *not be* demonstrated! This color change is dramatic and students should see it for the first time when they do the reaction.

## PROBLEM

To determine the amount of iron in a food sample by colorimetric analysis.

### Materials

Samples of food containing iron
(dates, raisins, leafy vegetables such as spinach, breakfast cereal)
Small test tubes
Evaporating dish
Beakers
10-mL pipet
100-mL graduated cylinder
Ringstand
Bunsen burner
$FeCl_3$
.1 *M* KSCN

● Encourage students to bring in cereal boxes or other containers listing the percent of iron in various food samples.

● Vitamin tablets containing iron can be used for testing. They must first be soaked in water to dissolve the colored coating. Then they can be dried and ground to a powder for analysis.

## PROCEDURE

### Part A: Preparing the Standards

● For this, and any other colorimetric experiment to work, great care must be taken in making dilutions.

1. Add 1.7 g of $FeCl_3$ (hexahydrate) to 99 mL of distilled water in a 250-mL beaker.
2. Mix the iron compound thoroughly until all of the solid is dissolved. Label this beaker 1%.
3. Carefully pipet out exactly 10 mL of this solution and place it in a second 250-mL beaker. Add 90 mL of distilled water and mix thoroughly. Label this beaker ".1%." Notice that it has been diluted 10 times.
4. Prepare a ".01%" solution, pipetting 10 mL of the ".1%" solution into a third beaker and adding 90 mL of water. Prepare and label beakers with concentrations of .01%, .001%, .0001%, .000 01%, and .000 001%.

● $FeCl_3$ is usually available as the hexahydrate:

$$FeCl_3 \cdot 6H_2O$$

1.7g of the hydrate in 99 mL of $H_2O$ will give a 1% solution.

5. Set up 6 small test tubes in a test tube rack. Label the tubes .1, .01, .001, and so on as you did with the beakers.
6. Place 5 mL (with a pipet) of each dilution from the appropriate beaker into the respective test tube. Start with the most dilute solution to prevent changing the concentration by accidentally transferring solution left in the pipet.
7. Add 5 mL of 0.1 *M* KSCN to each tube. Note the colors produced in each tube.
8. The tube labeled .000 001% should be a very faint pink, or colorless; if not, prepare another dilution from the .000 001% beaker and add this as your last standard.

● After the standards have been prepared, it is very easy to "run" several samples of various foods. You might assign several foods to each of several groups of students and pool their results.

### Part B: Testing for Iron

1. Raisins are easy to use for this experiment. Determine the mass of a sample of raisins. Record the mass in Table 23-1. Be sure the raisins are dry before determining the mass.
2. Place the sample in an evaporating dish (or crucible) and heat it strongly until the sample burns and only a gray ash remains.
3. After the dish has cooled, add 5 mL of distilled water to the dish and dissolve the remaining ash. Filter this solution. Transfer the filtrate to a test tube that is identical to those used to hold the standards.
4. Add 5 mL of KSCN solution to the test tube containing the sample.
5. Compare the color produced with that of the standard tubes. Match the color with the closest color in the standard tubes. This represents the % iron (as $Fe^{+3}$) in the sample. Record this value in Table 23-1.
6. Repeat the procedure for the other samples assigned to you.

● Caution students to avoid spattering when heating the sample.

**TABLE 23-1 • Iron in Food**

| | |
|---|---|
| Mass of Sample | |
| % of Iron | |
| Mass of Sample | |
| % Iron | |

## QUESTIONS

1. What is the approximate mass of iron in your sample of raisins?

*Answers will vary.*

2. The standards are made from a pure iron compound. The samples used for testing are not pure iron compounds. How will this affect your results?

*Results should be higher than normal.*

EXPERIMENT 24

Chapters 6, 7 Exc. 11

# Potassium in Coffee and Tobacco

Potassium was isolated from potash and identified by Sir Humphrey Davy in 1807. You won't isolate potassium from potash, but it can be easily isolated from and identified as a natural constituent of coffee and tobacco. You will isolate potassium and identify it by the characteristic yellow-colored complex that it forms with certain ions.

## PROBLEM

To identify potassium in coffee and tobacco.

### Materials

Coffee
Tobacco
Decolorizing carbon
1 *M* $KNO_3$
1 *M* $HNO_3$
Sodium hexanitrocobaltate solution
Filter paper
Funnel
Test tubes
Beaker
Bunsen burner
Stirring rods
Test tube holder

● Prepare the coffee (stale or fresh) for the students to use. This is not a quantitative test, so the volume of coffee tested need not be recorded.

● The formula for sodium hexanitrocobaltate is:
$Na[Co(NO_2)_6]$

## PROCEDURE

1. Place between 5 and 10 mL of coffee in a test tube. Add 0.1 g of decolorizing carbon and shake the tube for at least 1 minute.
2. Filter the contents of the tube, collecting the filtrate in another test tube.
3. Note the color of the filtrate. If it is not light yellow, add more carbon, shake, and filter again. Pour 3 to 5 mL of the filtrate into another test tube.
4. While gently agitating the tube, boil the filtrate until there is approximately 1 mL left.
5. Place the test tube in a beaker of tap water to cool the concentrated filtrate to room temperature.
6. Prepare three tubes, one containing a known sample of potassium, another with your sample to be tested, and a third with a blank—no potassium.

| | | |
|---|---|---|
| Tube 1: | Known K: | Add 1 drop of 1 *M* $KNO_3$ + 1 mL of distilled water. |
| Tube 2: | Sample: | Filtrate from step 5 |
| Tube 3: | Blank: | Add 1 mL of distilled water. |

7. To each of the three tubes, add 3 drops of 1 *M* $HNO_3$.
8. To each of the three tubes, add 10 drops of sodium hexanitrocobaltate solution.

● Caution students about heating liquids in a test tube—heat the sides of the tube while moving the tube through the burner flame.

9. Mix each tube with a separate, clean glass rod (or wood splint), and let them stand undisturbed for 3 to 5 minutes.

10. What change is evident in tube 1? What change is observed in tube 3? What change occurs in the sample tube?

*A yellow precipitate, $K_2Na[Co(NO_2)_6]$, a complex ion, forms in Tube 1.*

*This indicates the presence of $K^+$. Tube 3 should show no precipitate.*

*The sample should show a yellow precipitate.*

● If the colored liquid is carefully poured off, the colored precipitate can be seen.

11. Repeat the procedure using a small sample of tobacco in 5 to 10 mL of water.

## QUESTIONS

● If you wish to tell your students more about complex ions, consult any college-level chemistry test.

1. Describe the precipitate obtained with the coffee.

*Yellow*

2. Describe the precipitate obtained with tobacco.

*Yellow*

3. Place a small amount of KCl in an evaporating dish and heat it directly with a Bunsen burner. Look for a colored flame. Repeat the procedure with another salt. Is a flame of the same color produced?

*A violet flame should be produced with K salts.*

● If the violet flame is masked by Na, observe the flame through cobalt glass.

4. Repeat the test with a small bit of the tobacco and coffee sample. What color flame results?

*Violet*

5. How is the potassium ion, $K^+$, important in the human body? (You may wish to consult your school library.)

*$K^+$ is an important electrolyte in the body. It is also necessary*

*for proper nerve and muscle function.*

## EXPERIMENT 25

Chapters 6, 16

# Investigating Hard and Soft Water

● This experiment will give the students experience in ionization, precipitation, and solubility, and can serve as an entry into discussions of water purification and water pollution.

Water containing calcium and magnesium ions is said to be "hard." It has been observed for perhaps thousands of years that soaps lose their ability to form suds in hard water. Now there are several ways to remove the bothersome ions and make hard water "soft." There are also commercially available water softeners. Most city water supplies are treated to reduce the hardness in the water.

● If you wish to do Experiment 54, "Preparing Soap," first, you could use the soap prepared in that experiment. You might also do Experiment 55, "Preparing Dreft: A Detergent," at this point.

Although hardness is due mainly to the presence of calcium and magnesium ions, other elements such as strontium, barium, aluminum, manganese, iron, zinc, lead, and copper can also cause hardness. Fortunately, these metal ions are present only in trace amounts.

## PROBLEM

To show that water may contain minerals that make it hard and unfit for laundry purposes. To show that these ions can be easily removed.

● Ask students to check with the local water department to find out what your city does to reduce hardness in the water supply. Ask them to find out what commercial water softeners are available.

### Materials

Calcium chloride (solid)
5% $CaCl_2$
5% $MgCl_2$
Sodium carbonate
Calcium sulfate
Soap (solid and liquid)
Test tubes
Hard water

● Make hard water by dissolving 5 to 10 g of $MgSO_4$ in a liter of water.

## PROCEDURE

1. Arrange three test tubes in a test-tube rack.
2. Half fill the first tube with distilled water. Half fill the second tube with tap water. Half fill the third tube with *hard* water.
3. Add a pea-sized sample of solid soap to each tube and, with your thumb over the top, shake them well. Record your observations for each tube.

*Suds should form in the distilled water and tap water. Fewer suds should appear in the hard water.*

4. Half fill three test tubes with tap water. Add a small chip of soap to each.
5. To the first tube add 5 drops of 5% $CaCl_2$ solution. To the second tube add 5 drops of 5% $MgCl_2$ solution. Do not add anything to the third tube.
6. Shake each tube with your thumb over the mouth of the tube. Record your observations.

● Tube 3 is a control. Emphasize to the students the need for a control in an experiment such as this.

*Fewer suds will form in the tubes containing Ca and Mg.*

7. Half fill two test tubes with tap water. Add 0.5 g of calcium sulfate to each.

● Encourage students to express these reactions as ionic and net ionic reactions.
$Na_2CO_3$ (soluble) + $Ca^{+2}$ → $CaCO_3$ (insoluble) + $Na^+$

**8.** To the first tube add a few drops of soap solution and shake the tube. To the second tube add 0.5 g of sodium carbonate. Shake the tube and then add a few drops of soap solution. Shake the tube again. Record and explain your observations.

*Few suds in the first tube. A precipitate ($CaCO_3$) and suds form in the second tube.*

● Although we use the amount of suds as an indication, suds have nothing to do with the cleansing property of soap.

## QUESTIONS

**1.** Based on your observations, what is the difference between hard and soft water? How can you change hard water into soft water?

*Hard water contains Ca and Mg ions. If these are precipitated, the water becomes soft.*

**2.** What practical application is there for softening water?

*Since hard water prevents sudsing, washing and laundry work become difficult.*

**EXPERIMENT 26** Chapters 6, 7 Exc. 11

# Calcium and the Coagulation of Milk

The coagulation of milk involves a specific enzyme, rennin, which irreversibly converts a soluble protein into an insoluble coagulum. Calcium ions must be present for activation of this enzyme. Rennin converts the milk protein caseinogen into paracasein. Paracasein then combines with calcium ions to form an insoluble compound, calcium paracasein. Upon standing, this insoluble compound or coagulum contracts and squeezes out a clear liquid called whey.

In this experiment you will see the effects of calcium ion in the coagulation process. You can show that when calcium ions are removed from milk, coagulation is suppressed. When calcium ions are restored, coagulation occurs.

Coagulation of blood involves a similar process involving calcium ions. The blood clot squeezes out a clear fluid called serum. Some anticoagulants act by combining with the calcium ion, thus preventing it from reacting to form a blood clot.

● Sodium citrate or oxalate is added to fresh blood when it is collected to prevent it from clotting.
$Na_2C_2O_4$ (soluble sodium oxalate) + $Ca^{+2} \rightarrow$ $CaC_2O_4$ (insoluble calcium oxalate) + $Na^+$

## PROBLEM

To investigate the role of calcium in the coagulation of milk.

### Materials

| | |
|---|---|
| Fresh milk | Test tubes and stoppers |
| Rennin (Rennet) | Droppers |
| Sodium citrate | Water bath |
| 1% $CaCl_2$ | Thermometer |

## PROCEDURE

1. Add fresh milk to three test tubes until they are ¾ full and label the tubes 1, 2, and 3.
2. Place all three tubes in a water bath at 37°C.
3. To tube 1 add a small pinch of rennet. To tube 2 add a pinch of rennet and a pinch of sodium citrate. Tube 3 is a control. Do not add anything to it. Stopper the tubes.
4. After 5 minutes examine the tubes. Invert them, but do not shake the tubes. Record your results.

*Tube 1 coagulates, but tube 2 does not.*

● Tube 2 will not coagulate because sodium citrate reacts with the calcium ion to produce insoluble calcium citrate. This prevents $Ca^{+2}$ from causing coagulation.

5. Add 4 to 5 drops of calcium chloride solution to tube 2. Replace it in the water bath for 5 minutes. Record your results.

*Coagulation occurs.*

● When calcium chloride is added and tube 2 heated, the milk will coagulate because of the presence of $Ca^{+2}$.

## QUESTIONS

**1.** What happened to the calcium ion in tube 2?

*It is precipitated as calcium citrate.*

**2.** Explain what happened when $CaCl_2$ was added to tube 2.

*Coagulation occurred because $Ca^{+2}$ was added to the milk.*

EXPERIMENT 27

Chapters 10, 20

# Chemistry of the Permanent Wave

● Ask your local beautician to save hair for your class—especially long hair.

Hair is a fibrous protein. Proteins are composed of long chains of amino acids. These chains are called polypeptides. Hair belongs to the group of fibrous proteins called alpha-keratins. A single hair is composed of long strands of alpha-keratin molecules. Electron microscopy has shown that these protein strands form fibrils. The fibrils, usually 6 or 8, twist around each other, much like strands in a rope, to form hair.

Alpha-keratins have two specific characteristics that help explain some of the properties of hair. First, they stretch when heated. Normal hair will stretch to almost double its length when moistened and heated. After cooling, it returns to its normal length. Secondly, hair contains a specific amino acid called cystine. This particular amino acid is unusual because it has a disulfide part that can act as a bridge between polypeptides, holding them together. This bridge is a covalent bond between sulfur atoms of two different polypeptide chains. This is what gives hair its shape.

Normal hair is composed of 10 to 14% cystine. If the disulfide bonds in the cystine molecules are broken, the hair can be reshaped. Later, new disulfide bridges will form and the hair will stay in its new shape.

Most permanent wave lotions contain a compound that breaks the disulfide bridges between polypeptides and causes hair to lose its shape. In this experiment, you will investigate this reaction. You will treat hair with a solution that breaks the disulfide bond, and test for the presence of the sulfhydryl (SH) group formed.

## PROBLEM

To study the chemistry involved in the permanent wave.

### Materials

Hair
Permanent wave lotion
Ammonium nitrate (solid)
NaOH (concentrated)
2% sodium nitroprusside
Test tubes
Glass rod
Rubber bands

● Permanent wave lotion often contains thioglycolic acid, which acts as a reducing agent. This ruptures the disulfide linkages in hair.

## PROCEDURE

### Part A: Determining the Presence of the Sulfhydryl (SH) Group

1. Put a pinch of hair in each of three test tubes. Label the tubes 1, 2, and 3.
2. Add ½ tube of distilled water to tube 1. Add ½ tube of permanent wave lotion to tubes 2 and 3.
3. Allow the hair to remain in the tubes for 20 minutes. After that time, wash the hair in tubes 2 and 3 until the odor of the lotion is gone.

4. To test for the presence of the sulfhydryl group, add 5 mL of water and 2 g of solid ammonium nitrate to tubes 1 and 2. Shake the tubes until all of the solid has dissolved. Add 2 drops sodium nitroprusside solution and 2 mL of concentrated NaOH to tubes 1 and 2. A purple color indicates the presence of the sulfhydryl group. Record your results in Table 27-1.
5. Check the container of the permanent wave lotion used. If it is self-neutralizing, allow tube 3 to stand overnight and test as you did in step 4. If the material you used has a separate neutralizer, add the neutralizer to tube 3 and allow it to react for 20 minutes. Then, test for sulfhydryl as you did in step 4. Record your results.

● The neutralizer is a mild oxidizing agent, such as hydrogen peroxide ($H_2O_2$), which permits the formation of disulfide bonds.

**TABLE 27-1 • Testing Hair**

| TEST TUBE | POSITIVE FOR (SH) | NEGATIVE FOR (SH) |
|---|---|---|
| 1 | | x |
| 2 | x | |
| 3 | | x |

### Part B: Waving Hair

1. Select a length of long, straight hair. Wind it around a glass stirring rod or a popsicle stick.
2. Hold the ends down with rubber bands.
3. Apply the lotion, rinse, and add the neutralizer according to the directions on the container of wave lotion you are using.
4. Describe the results.

*A winding wave should result.*

## QUESTIONS

1. Explain the effect of the neutralizer on tube 3.

*The neutralizer permits the formation of disulfide bonds.*

2. Explain why some people's hair is more wavy than others.

*Variations in the amount of cystine.*

EXPERIMENT 28

Chapters 13, 15

# Heat Effects in Simple Reactions

Be sure students understand the meanings of exothermic and endothermic.

Some substances generate heat when placed in solution while others absorb heat. A common example of a substance that generates heat when placed in solution is sulfuric acid. However, it must be done cautiously, since the danger of spattering is great. NaOH in solution also produces heat. Here again, care must be exercised in handling solid NaOH.

In this experiment you will examine two reactions, one that is exothermic and one that is endothermic. Although the qualitative heat effects can be easily determined, you are asked to determine quantitatively the change in heat ($\Delta H$) for these reactions. You will also observe reactions in which two compounds are combined to produce or absorb heat. These heats of reaction can be determined by measuring the extent to which the reaction changes the temperature of a measured amount of water.

## PROBLEM

To investigate exothermic and endothermic reactions.

### Materials

Calcium chloride (solid)
Laundry bleach
.5 *M* sodium sulfite
Barium hydroxide (solid)
Ammonium thiocyanide (solid)
Ammonium nitrate (solid)
Styrofoam cups
Small beakers
Thermometers
Stirring rod

.50 *M* sodium sulfite solution: 63.0 g $Na_2SO_3$/liter solution.

If styrofoam cups are used, the change in temperature will be much greater. However, students can better feel the increase or decrease in temperature if small beakers are used.

## PROCEDURE

### Part A: Heat of Solution

1. Place 100 mL of water in a styrofoam cup or a beaker. Determine the temperature of the water and record it in Table 28-1.
2. Determine the mass of a 10 to 15 g sample of ammonium nitrate. Record the mass in Table 28-1.
3. Quickly dump the ammonium nitrate into the styrofoam cup and stir with a wooden splint or glass rod.
4. Record your observations. Again determine the temperature and record it in Table 28-1.

100 mL is used to make the calculation easier.

620 J (148 calories) are absorbed per 1 mL of water.

---

5. Place 100 mL of water in another styrofoam cup. Determine the temperature of the water and record it in Table 28-1.

The temperature should drop considerably (6-9°C).

6. Determine the mass of a 10 to 15 g sample of calcium chloride. Record the mass in Table 28-1.
7. Dump the calcium chloride into the cup and stir.
8. Determine and record the temperature and record your observations.

● The reaction is exothermic: 490 J (117 calories) are given off per mL of water. If calcium chloride is not available, you can also use barium oxide. The temperature should increase about 12°C.

**TABLE 28-1 • Heat of Solution**

| | |
|---|---|
| Temp. of water (step 1) | |
| Mass of Ammonium Nitrate | |
| Temp. of Water (step 4) | |
| Temp. of Water (step 5) | |
| Mass of Calcium Chloride | |
| Temp. of Water (step 8) | |

## Part B: Heat of Reaction

1. Add 50 mL of bleach to a styrofoam cup. Determine and record the temperature of the bleach in Table 28-2.
2. Add 50 mL of a .50 *M* sodium sulfite solution to the bleach.
3. Record your observations. Determine and record the temperature of the solution.

● The temperature should increase about 20°C.

4. Place 16 g of barium hydroxide in a 100-mL beaker.
5. Add 8 g of ammonium thiocyanide to the beaker.
6. Stir the two solids together with a wooden splint.
7. Record your observations.

● This illustrates the fact that heats of reaction can occur without solutions!

8. Make a small puddle of water on your laboratory desk. Put the beaker on the film of water. Record your results.

● The reaction is:
$Ba(OH)_2 \cdot 8H_2O\ (s) + 2NH_4SCN\ (s) \rightleftharpoons Ba(SCN)_2\ (s) + 2NH_3\ (g) + 10H_2O\ (l)$

*The reaction is so endothermic that the beaker will freeze*

*to the table top.*

**TABLE 28-2 • Heat of Reaction**

| | |
|---|---|
| Temp. of Bleach | |
| Temp. of Bleach + Sodium Sulfite | |

## QUESTIONS

1. When ammonium nitrate is added to water is the reaction exothermic or endothermic?

*Endothermic*

2. Calculate the $\Delta H$ for the reaction (per mole of ammonium nitrate).

*Mass of $NH_4NO_3$/80 g $NH_4NO_3$ per mole = moles of $NH_4NO_3$;*

*(Temperature (initial) − Temperature (final) x 100 = calories of heat*

*absorbed; calories of heat absorbed/moles of $NH_4NO_3$ = $\Delta H$ (cal/mole*

*$NH_4NO_3$) convert to Kcal.*

3. When calcium chloride is added to water is the reaction exothermic or endothermic?

*Exothermic*

4. Calculate the $\Delta H$ for the reaction (per mole of calcium chloride).

*Same as #2 except $CaCl_2$ = 110 g/mole*

5. When bleach and sodium sulfite are mixed, is the reaction exothermic or endothermic?

*Exothermic*

6. Calculate the $\Delta H$ for the reaction per mole of NaHOCl. Assume that bleach is .7 *M*.

*Approximately −80 Kcal/mole*

7. Is the reaction between barium hydroxide and ammonium thiocyanide exothermic or endothermic? What practical purposes might such a reaction have?

*Endothermic. Cold packs, first aid, and so on.*

● This experiment introduces the student to the use of enzymes as catalysts. It also shows an example of osmosis. You may wish to use this experiment with the biochemistry chapter since it illustrates the glycolysis pathway in bacteria. The sauerkraut made in this experiment is edible. You can also make dill pickles by the same reaction.

**EXPERIMENT 29** Chapters 15, 20

# Lactic Acid and Sauerkraut

Sauerkraut is characterized by its tangy, sour taste—not at all like that of the cabbage from which it is made. Pioneer families (and even some modern families) often made sauerkraut. You will prepare sauerkraut, using an old method, to show the production of lactic acid.

Lactic acid bacteria convert the sugar in cabbage into lactic acid in sauerkraut. Lactic acid is responsible for the sour taste in sauerkraut. Most of the home recipes for sauerkraut do not require the addition of water. This is because NaCl is added to help pull water and sugar out of the cabbage. The sugar can then be acted upon by the lactic acid bacteria normally present. This movement of water is an example of osmosis. The water leaves the cabbage to enter the more concentrated salt solution. In this experiment you will add both water and salt. You must take care, however, not to add too much salt, since excessive salt will inhibit the activity of the lactic acid bacteria.

## PROBLEM

To prepare sauerkraut by the utilization of enzymes from bacteria.

### Materials

Cabbage
NaCl
Universal pH paper
Jar with lid
(from home—peanut butter jar or similar type jar)
Beaker

● Ask students to bring in a jar several days prior to beginning the experiment.

## PROCEDURE

### Part A: Preparing the Sauerkraut

● Your school cafeteria will gladly shred a cabbage for you.

1. Shred or finely chop a cabbage, pulling off and discarding the outside leaves.
2. Pack the shredded cabbage into the jar until the jar is about ½ full.
3. Add about .5 g of NaCl, sprinkling it over the cabbage.
4. Continue to add the shredded cabbage to the jar until it is almost full.
5. Add another .5 g of NaCl over the top of the cabbage.
6. Slowly add water until the jar is full. Tap the jar to remove as many air bubbles as possible.
7. Measure the pH of the solution with pH paper and record it.

*pH approximately 7*

8. Cap the jar—but do not put the cap on tightly.

● Note that this reaction requires about two weeks.

9. Place the jar where it will not be disturbed for about two weeks. The jar should be at room temperature. Check occasionally to be sure it is covered with water. Add water when neccessary.

**Part B: Testing the Sauerkraut**

1. After two weeks, pour the solution from the sauerkraut into a beaker.
2. Test the pH of the solution with pH paper. Record the pH.

*pH should be acidic*

3. Taste the sauerkraut. If you wish, you can wash your sauerkraut, place it in a beaker of water, and bring the water to a boil. As soon as it begins to boil, remove the sauerkraut.

● You might like to follow this with Experiments 47 and 48 from this lab manual. Have students titrate the lactic acid ($HC_3H_5O_2$) to determine pH.

## QUESTIONS

1. How does your sauerkraut compare with that which you buy at the grocery store?

*It will taste the same. However, commercial sauerkraut is bleached to give it a white appearance.*

2. Devise your own experiment for preparing *pickles*.

● This experiment can easily be performed at home. Encourage students to try some of these experiments at home.

**EXPERIMENT 30**

Chapters 15, 20

# Plant Enzymes That Digest Protein

Meat tenderizer is often sprinkled on meat prior to cooking. Most meat tenderizers contain, as the active ingredient, the enzyme papain. Papain is a proteolytic enzyme that is extracted from unripe papaya, a tropical melon. This enzyme catalyzes the hydrolysis of protein to amino acids. Unlike most enzymes, which are pH specific, papain is effective over a pH range of 4 to 8. There are other similar proteolytic enzymes that are derived from plants. Ficin is produced by the fig family. Bromelin is found in pineapple. Most proteolytic enzymes are of animal origin and are not effective when synthesized.

In this experiment you will study the effect of proteolytic enzymes on protein. In order to better observe the hydrolysis process, you will use gelatin as a protein source.

## PROBLEM

To study the effect of proteolytic enzymes on protein.

### Materials

| | |
|---|---|
| Meat tenderizer | Trough |
| Gelatin | Ice |
| Evaporating dishes | Pineapple (or figs) |
| Beakers | |

## PROCEDURE

### Part A: Papain

1. Prepare some gelatin by following the directions on the package. However, prior to cooling the gelatin, pour equal portions of the liquid into two evaporating dishes. Label the dishes 1 and 2.
2. Generously sprinkle meat tenderizer on top of the hot gelatin solution in dish 1.
3. Solidify the gelatin in both dishes by placing them in an ice bath (a large trough filled with ice).
4. After the gelatin has become firm, sprinkle meat tenderizer generously over the top of dish 2.
5. Allow the gelatin to sit at room temperature for a few minutes. Note any changes in the gelatin. Gently poke it with your pencil or a glass rod. Record your observations.

*The hot gelatin inactivates the papain and the gelatin solidifies.*

*However, the gelatin in dish 2 will have a liquid on top, due to the action of papain.*

**Part B: Bromelin**

1. Bring a fresh pineapple from the grocery store.
2. Prepare a dish of gelatin as you did in Part A.
3. When the gelatin is solid, place a slice of fresh pineapple, or fresh pineapple juice, on top of the gelatin. Record your observations.

*The surface should liquefy.*

● You can prepare an extract of ficin or bromelin from fresh figs or pineapple. Simply cut these and squeeze the juice. Filter to remove the pulp.

## QUESTIONS

1. What effect does the enzyme have on the protein in gelatin?

*It breaks it down, or hydrolyzes it.*

2. What effect does heat have on enzyme activity?

*Since an enzyme is a protein, heat coagulates it.*

3. What effect does the pineapple have on the protein in gelatin?

*It breaks it down.*

4. Do you think you could make a congealed salad with gelatin and fresh pineapple?

*No, because the gelatin would partially liquefy.*

5. Why can you make a congealed salad with gelatin and canned pineapple?

*Canning destroys the enzyme.*

6. Some people use meat tenderizer while they are grilling or barbecuing meat. Is this effective?

*Probably not, since the heat inactivates the enzyme.*

EXPERIMENT 31

Chapters 15, 20

# Enzymes in Typical Foods

Most foods contain a variety of enzymes. Enzymes catalyze metabolic reactions in the body. In this experiment you will test foods for the presence of the enzyme catalase. Catalase catalyzes the decomposition of hydrogen peroxide into water and oxygen gas.

$$2H_2O_2 \xrightarrow{\text{catalase}} 2H_2O + O_2\,(g)$$

Perhaps you have noticed that when hydrogen peroxide is placed on a cut or an open wound, fizzing and bubbling occur. This is caused by catalase in the blood decomposing the peroxide and releasing oxygen gas. Catalase is found in many foods. It is especially concentrated in turnips and rutabaga. In addition to determining which foods contain this specific enzyme you will also check to see what effect cooking has on the activity of catalase.

● Enzyme activity is often measured in turnover numbers, that is, how many molecules of substrate can be acted upon by one enzyme molecule per minute. The turnover number for catalase (a *slow* enzyme) is 5 million! Since enzymes are substrate specific, catalase does nothing except catalyze this particular reaction.

## PROBLEM

To determine which foods contain the enzyme catalase.

● Obtain samples of fresh and cooked foods of the same variety. Solid food should be crushed.

### Materials

3% $H_2O_2$ (drugstore variety)
Raw turnips
Raw rutabaga
Raw potatoes
Several samples of raw foods
Several samples of cooked foods
Beakers or paper cups
Graduated cylinders

## PROCEDURE

1. Pour 10 to 15 mL of $H_2O_2$ into several small beakers or paper cups.
2. Add a different sample of food to each container.
3. Record your observations for each food in Table 31-1.

● Students should look specifically for bubbles of gas, evidence that oxygen is produced by the breakdown of hydrogen peroxide.

**TABLE 31-1 • Catalase**

| FOOD | OBSERVATIONS |
|---|---|
| | |
| | |
| | |
| | |
| | |
| | |

## QUESTIONS

1. What did you observe when turnips were added to $H_2O_2$?

*A gas (oxygen) was evolved.*

2. Compare the rates of reaction in the various containers.

*Answers will vary.*

3. How does the reaction of cooked foods compare with the reaction of raw foods? Explain your answer.

*Cooking destroys the enzymes, therefore no reactions occur.*

# Making Cheese

Cheese is the solid portion of milk that has been separated from the liquid portion. This solid portion is called the curd and the remaining liquid portion is called the whey. An enzyme is necessary to accomplish this separation, otherwise production of cheese would be too slow to be practical. This enzyme is rennet. Rennet is now produced commercially, but at one time it could only be extracted from the fourth stomach of the suckling calf.

● Rennet can be purchased in grocery stores, under the trade name *Junket*.

There are basically three steps in preparing cheese from milk.

1. **Preparation for Curdling**
   Milk contains lactose, often called milk sugar. Lactose is converted into lactic acid when milk sours. It is the lactic acid that gives sour milk its characteristic taste. The first step in making cheese is to increase the lactic acid content of the milk by converting as much lactose to lactic acid as possible. Commercially, bacteria are used for this purpose. In this experiment, however, buttermilk will be added as a "starter" several hours before you begin the investigation.

2. **Curdling**
   Curdling occurs when the enzyme rennet is added.

3. **Aging (or Curing)**
   There are many techniques used in the aging process. It is during this time that the cheese develops its characteristic flavor. Cheese is usually aged between a temperature of 35 and 55°C. Aging is a chemical process and like all chemical reactions, it is accelerated at higher temperatures. Aging may take a couple of weeks to a couple of months.

● The enzyme rennilase also works well. Rennilase can be purchased from biological supply companies.

## PROBLEM

To make cheese by the action of the enzyme rennet.

### Materials

● Prepare the milk sample by adding 1.0 mL of buttermilk per 200 mL of whole milk and let it sit for about 4 hours before using it in this experiment.

Milk
Rennet or rennilase
Beaker
Thermometer
Graduated cylinder
Droppers
Bunsen burner
Stirrer
Filter paper
Funnel
Ringstand

## PROCEDURE

1. Place 200 mL of milk in a large beaker.
2. Warm the milk to 30 to 32°C. Be careful *not* to overheat! Check the temperature with a thermometer. (Do not stir the milk with the thermometer.)
3. While stirring add ¼ to ½ of a rennet tablet (or 5 drops of rennilase) to the milk. Use a stirring rod to break up the curd that forms into tiny bits.
4. Gradually increase the temperature to 38°C, but no higher!

5. Remove the beaker from the flame and allow it to remain undisturbed for at least 15 minutes. Observe what happens.
6. Filter the contents of the beaker. Your cheese is on the filter paper.
7. Press all of the liquid from the cheese and allow it to dry.

## QUESTIONS

1. How does your cheese differ from commercially prepared cheese in appearance and taste?

*The students' cheese will be soft and white (or gray). It will taste bland.*

*(Commercially, cheese is salted after step 7.)*

2. What is the *curd*? What is the *whey*?

*Curd is the solid part of milk; whey is the liquid remaining after the solid is removed.*

3. Name some types of cheeses. How do you think these are made?

*Encourage students to go to the library or write for information. Try the State Dairy Council, or U.S. Dept. of Agriculture.*

EXPERIMENT 33 Chapters 15, 20

# The Chemistry of Bruised Fruit

It is a common observation that apples, pears, bananas, and other fruits darken when they are bruised or peeled and exposed to air. This darkening is caused by enzymes released from the damaged cells of the fruit. The enzymes are oxidative enzymes. They react with oxygen to digest the cells of the fruit. This reaction with oxygen causes rapid changes in the color and taste of the fruit.

● It is now believed that the same enzyme responsible for the darkening of fruit is responsible for the formation of melanin—the pigment in skin which gives it a dark color.

In this experiment you will investigate the effect that oxygen has on the darkening of bruised fruit. You will also test the effect of vitamin C in retarding the oxidation reaction between the enzymes and oxygen. Vitamin C is a naturally occurring reducing agent found in fresh fruits.

## PROBLEM

To investigate the effect of oxygen on bruising of fresh fruit.

### Materials

Fresh fruit (preferably apples)
Vitamin C tablets
Fruit juice (orange, tomato)
Test tubes with stoppers

● The 100-mg vitamin C tablets will suffice.

## PROCEDURE

● Have a jar of boiled tap water prepared for use by the class.

1. Before preparing the apple for this experiment, it is necessary to set up 6 test tubes with appropriate solutions.
   Tube 1: Empty
   Tube 2: Empty
   Tube 3: ½ filled with tap water
   Tube 4: Completely filled with water that has been boiled
   Tube 5: ½ filled with a solution made by dissolving 1 vitamin C tablet in a small amount of water
   Tube 6: ½ filled with fruit juice

● If you don't have vitamin C tablets, you can use any powdered fruit juice product which has a very high vitamin C content.

2. Cut an unpeeled apple into six or more equal pieces. Each piece must be small enough to be completely submerged in the tubes ½ filled of liquid.
3. Using your thumb, apply just enough pressure to bruise each piece.
4. Immediately place a bruised slice of apple in each test tube. Stopper all tubes except tube 1.
5. After 20 to 25 minutes, remove and examine each apple slice. Using the amount of discoloration in tube 1 as a standard, record the relative amount of discoloration in each tube.

● Encourage students to organize their own methods to compare colors and collect data in this experiment.

## QUESTIONS

1. How did the slice of bruised apple in tube 1 compare with the slices in the other tubes?

*It should be much darker in color.*

2. Is there evidence that oxygen is involved in the reaction causing fruit to bruise?

*Yes, because of tube 1's darker color.*

3. Vitamin C is a reducing agent. Is there evidence that vitamin C can retard bruising in fruit? Explain how this occurs.

*Yes. Essentially, vitamin C reacts with the enzyme before it can react with the apple or fruit.*

4. Is there a difference between the effect of natural vitamin C (fruit juice) and artificial vitamin C (vitamin tablet)?

*Very little can be noted in this experiment.*

5. How does nature prevent fruit from reacting with oxygen?

*Skin, peel, and so on.*

6. When preparing a fruit salad that includes sliced or chopped apples, cooks often squeeze a lemon or lime over the salad. How does this prevent the fruit from darkening?

*Vitamin C in lemon juice serves the same function as the vitamin C in the tablet does in this experiment.*

● If students show interest in vitamin C during this experiment, you might like to have them do Experiment 50, "Testing For Vitamin C in Foods," at this time.

● Substances such as vitamin C which prevent oxidation are called anti-oxidants.

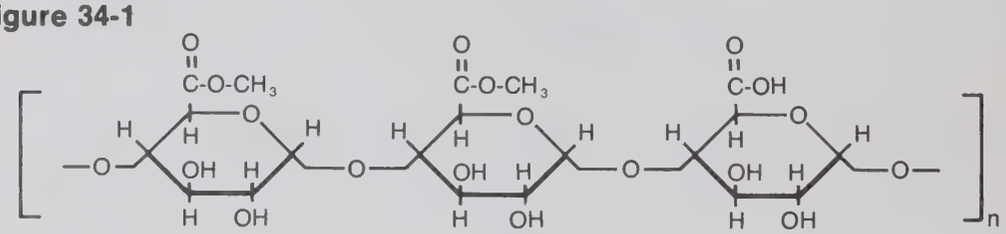
EXPERIMENT 34

Chapters 15, 20

# Pectin in Fruit Juices

Pectin is a very large, colloidal molecule found in all fruits and fruit juices. The structure of pectin is shown in Figure 34-1. Notice the repeating units in the structure. These units are glacturonic acid molecules held together by oxygen linkages. Pectin has a protective function in fruits. It is primarily responsible for preventing particles from settling out in the fruit. Pectin will lose its colloidal property if it decomposes by reacting with water. Ordinarily, the reaction with water is very slow. However the enzyme pectinase will catalyze the reaction to decompose pectin into individual sugars. Thus, to prepare fruit juices from fruits, pectinase is added to the pulp. This causes the breakdown of the large pectin colloid into smaller sugar units. This, in turn, allows the fruit juice to be released from the cells of the fruit. In this experiment you will observe the action of pectinase in preparing apple juice from applesauce.

**Figure 34-1**




## PROBLEM

To observe the action of the enzyme pectinase on the breakdown of fruit pectin.

● The enzyme pectinase can be purchased from biological supply companies.

### Materials

Pectinase
Applesauce
Filter paper (No. 201 works best)
Beakers
Graduated cylinders
Funnels
Stirring rods

## PROCEDURE

1. Obtain two 100-mL graduated cylinders and two 250-mL beakers. Label one beaker and cylinder "A"; label the other beaker and cylinder "B."
2. To beaker A add 50 g of applesauce and 1 mL of pectinase. Stir thoroughly.
3. To beaker B add 50 g of applesauce but no enzyme.
4. Let both beakers stand for 10 minutes.
5. Place a funnel into each of the 100-mL graduated cylinders and fold a piece of filter paper for each.
6. Filter solution A into graduated cylinder A, and solution B into graduated cylinder B.

7. By reading the volume directly, record the amount of apple juice obtained from beaker A and the amount obtained from beaker B.

*Beaker A* *Beaker B*

## QUESTIONS

1. What practical application does the addition of pectinase to pulp have?

*Maximum production of fruit juice in industry.*

2. Can you explain why most fruits do not readily release their juices?

*The enzyme pectinase is not ordinarily active in natural fruit. (Students may answer that some fruit does not contain pectinase.)*

● You might like to look at concentration of enzyme versus amount of juice produced. Assign student teams to work with no enzyme, 0.25 mL, 0.50 mL, and 0.75 mL of pectinase and plot these concentrations versus the amount of fruit juice each produced. Try other fruit pulps.

● This is a quick and easy experiment that can be used to illustrate chemical reactions, reaction of ions in solution, and salt formation. Solubility is also demonstrated in this experiment.

**EXPERIMENT 35**

Chapters 10, 16

# Making Table Salt

Table salt is a product of the reaction of an acid (HCl) and a base (NaOH). A salt is a general product of the reaction of an acid and a base. Although we usually think of sodium chloride when we think of salt, the term salt can refer to many different compounds. Sodium chloride is an ionic solid consisting of positive sodium ions and negative chloride ions. There are no molecules of NaCl. Most chlorides are soluble in water. However, silver chloride, AgCl, is insoluble in water. This provides the basis for identifying the $Cl^-$ ion.

## PROBLEM

To prepare sodium chloride (table salt) and test for the presence of the chloride ion by precipitating it as AgCl.

### Materials

● 0.1 *M* $AgNO_3$—17 g $AgNO_3$/L

| | |
|---|---|
| 3 *M* $HNO_3$ | Glass stirring rod |
| 0.1 *M* $AgNO_3$ | Blue litmus paper |
| 1.0 *M* HCl | Test tube |
| 1.0 *M* NaOH | Bunsen burner |
| Evaporating dish | Wire gauze |
| Graduated cylinder | Ringstand |
| Droppers | Dropper |

## PROCEDURE

### Part A: Making the Salt

1. Add 5.0 mL of NaOH to the evaporating dish.
2. Drop in a strip of blue litmus paper.
3. While stirring constantly with a glass rod, slowly add HCl to the dish until the litmus paper just begins to turn pink.
4. Again while stirring, add NaOH with a dropper until the paper is neither pink nor blue—somewhat violet in color.
5. Place the evaporating dish on a wire gauze and heat it slowly until the liquid completely evaporates.
6. Describe the solid in the evaporating dish.

*A white, crystalline material*

### Part B: Testing the Salt

**CAUTION** $AgNO_3$ will stain the skin and clothing.

1. Scrape off a small amount of the NaCl and add it to a test tube. Dissolve it in about 20 drops of distilled water.
2. Add 2 drops of 3 *M* $HNO_3$.
3. Add 3 to 4 drops of 0.1 *M* $AgNO_3$ and mix with a stirrer.
4. Record your observations.

*A white, cloudy precipitate forms.*

## QUESTIONS

**1.** Describe the appearance of the solid left in the dish after evaporation.

*A white, crystalline solid*

**2.** Write the reaction for making salt.

$HCl(aq) + NaOH(aq) \rightarrow H_2O + NaCl(aq)$

**3.** Why is it necessary to use distilled water?

*Tap water probably contains $Cl^-$. (Check by adding $HNO_3$ and $AgNO_3$ to tap water.)*

**4.** Describe the precipitate obtained by adding $AgNO_3$.

*A cloudy, white precipitate*

**5.** Write the equation for the reaction of $AgNO_3$ and salt.

$Ag^+(aq) + Cl^-(aq) \rightleftharpoons AgCl(s)$

● This experiment shows the importance of colligative properties of solutes. One such property is the ability of a solute to lower the freezing point of water. The freezing point of water is lowered 1.86°C for each mole of solute species added per kilogram of water. There are *two* solute species in NaCl.

● The best way to illustrate these reactions is to *perform the experiment*! Borrow an ice-cream freezer and make a bucket of ice cream during the class.



# Salt and Homemade Ice Cream

When making homemade ice cream, rock salt is added to the ice in the bucket. Why? What effect does this have on the temperature of the ice-water mixture? Every winter in colder climates salt is placed on icy roads and sidewalks. This makes the ice melt. Why? In this experiment you will attempt to find the answers to these questions.

## PROBLEM

To determine the melting point of ice when salt or other chemicals are added to it.

### Materials

Sodium chloride
Calcium chloride
Calcium carbonate
Methyl alcohol
Mineral oil
Ethylene glycol
Ice
Styrofoam cups
Thermometer
Stirring rod

## PROCEDURE

### Part A: The Melting Point of Ice

1. Place some ice in a styrofoam cup.
2. Determine the melting point of the ice.

*It should be 0°C.*

3. Does the temperature change as the ice continues to melt?

*It should not.*

4. Does stirring the ice as it melts affect the melting point? (Don't stir with the thermometer!)

*No*

### Part B: Effects of Chemicals on Melting

1. Fill the styrofoam cup ¾ full of ice.
2. Add 10 g of sodium chloride and mix the ice and salt.
3. Determine the melting point and record your results.

*The ice should melt at a lower temperature.*

4. Based upon your observations with NaCl, predict what effect the chemicals listed in Table 36-1 would have on the melting point of ice.
5. Determine the melting point of an ice-10 g mixture for each chemical in Table 36-1. Record your results.

**TABLE 36-1 • Melting Ice**

| CHEMICAL | EFFECT ON MELTING POINT OF ICE (PREDICTION) | EFFECT ON MELTING POINT OF ICE (EXPERIMENTAL) |
|---|---|---|
| Calcium chloride (soluble in $H_2O$) | | *Lowers it* |
| Calcium carbonate (insoluble in $H_2O$) | | *No effect* |
| Methyl alcohol (soluble in $H_2O$) | | *Lowers it* |
| Ethylene glycol (soluble in $H_2O$) | | *Lowers it* |
| Mineral oil (insoluble in $H_2O$) | | *No effect* |

## QUESTIONS

**1.** What substances influence the melting point of ice?

*Calcium chloride, methyl alcohol, ethylene glycol*

**2.** What characteristics do they have in common?

*Solubility in water*

**3.** Why is salt added to the road to remove ice?

*If the temperature is, say –3°C, adding salt to an icy road will lower the freezing point of the salt solution formed to below –3°C. Thus, the surrounding temperature will be greater than the melting point, and the ice will melt.*

# Mayonnaise, an Edible Emulsion

Perhaps you have noticed that a mixture of oil and vinegar tends to separate into layers. If you shake an oil and vinegar mixture and observe it carefully, you will see tiny drops of oil. These droplets enlarge and eventually form a separate layer of oil. The oil droplets can be prevented from coming together, however, by adding a third substance that will stabilize the two immiscible liquids. This will prevent the liquids from separating into layers. The stabilizing substance is called an emulsifying agent, and the resulting mixture of the three substances is an emulsion. An emulsion is not a solution. It is a dispersion of one liquid in another liquid with an emulsifier for stabilization.

Soap is an example of an emulsifying agent. When greasy dishes are washed in soap (or detergent), the hot water changes the grease into oil. The detergent emulsifies the oil so that it is removed from the dishes by the water and washed down the drain. Mayonnaise is another example of an emulsion. In mayonnaise the emulsifying agent is egg yolk. In this experiment you will prepare mayonnaise by emulsifying salad oil and vinegar.

● The emulsifying agent in egg yolk is the phospholipid lecithin.

## PROBLEM

To prepare mayonnaise, a typical emulsion.

● Be sure that all ingredients are at room temperature.

### Materials

Eggs
Vinegar
Salt
Prepared mustard
Salad oil
Small bowl or large beaker
Electric beater or egg beater
Magnifying glass

● Let students work in pairs —one can add the materials while the other stirs or beats.

## PROCEDURE

1. Place 2 egg yolks, ¼ teaspoon of mustard, ½ teaspoon of salt, and 1 teaspoon of vinegar in a large beaker or bowl.
2. Beat the mixture until the egg yolks are sticky. Obtain 250 mL (1 cup) of salad oil.

● If the emulsion gets too thick, add another spoonful of vinegar.

3. Begin adding the salad oil, a small amount at a time, while beating constantly.
4. After adding about ⅓ cup of oil, the mixture should begin to get thick. When this happens, add the remainder of the oil while beating the mixture.
5. Continue beating until you have an even dispersion.

## QUESTIONS

1. Describe the mayonnaise you prepared.

*Yellow, thick, and glistening*

2. How does it differ from the commercial product?

*In color. It may also have a more oily appearance.*

Have a commercial variety of mayonnaise available for comparison.

3. Using a hand lens or a magnifying glass, examine a small sample of your mayonnaise. Describe it.

*Small oil globules can be observed.*

4. What other products can you think of that exist as emulsions?

*Milk, latex rubber, oil in water, etc.*

# Making Hand Lotion

Glycerol is commonly used as a humectant—a substance that helps retain moisture. In this experiment you will use it to make hand lotion. You will prepare an emulsion of water, glycerol, and paraffin. Cosmetics with a glycerol base evaporate and leave a film on the skin as they dry. This produces a cooling effect and reduces loss of moisture from the skin. The hand lotion you will make in this experiment is very similar to that which you would buy at the drugstore. You will prepare the hand lotion using two different formulas.

## PROBLEM

To prepare hand lotion, a humectant and an emulsion.

### Materials

Soft paraffin wax
White petrolatum
Liquid paraffin
Diglycol stearate
Bunsen burner
Polyethylene glycol
Ringstand
Glycerol
Thermometer
Anhydrous lanoiin (white)
Stirring rod
Boric acid
Perfume (inexpensive)
Beakers

## PROCEDURE

### Part A: Formula 1

1. In a 300-mL beaker, mix and melt the following substances.
   13.5 g of soft paraffin wax
   3.5 g of white petrolatum
   13.5 g of liquid paraffin
   13.5 g of diglycol stearate

● Demonstrate the proper set-up and precautions to be observed when using a water bath.

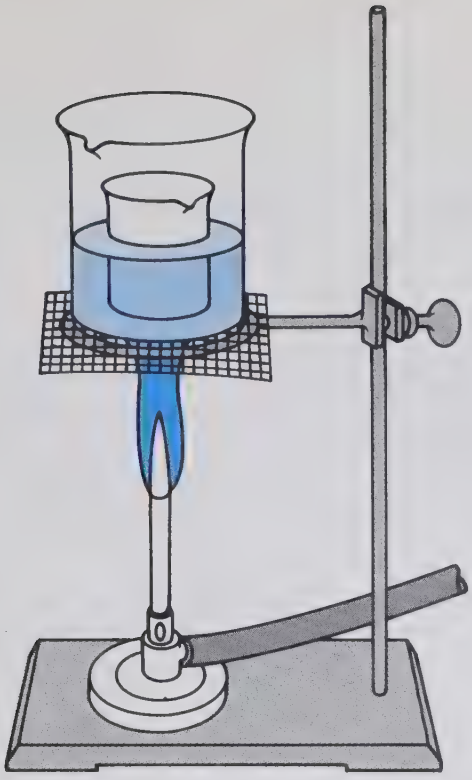
2. After the mixture has melted, heat it to 70°C, using a steam bath. (See Figure 38-1)
3. In a second beaker, add 10 mL of polyethylene glycol to 54 mL of water. Heat this solution to 70°C.
4. With constant stirring, add the water-glycol solution to the first solution. Remove the beaker from the water bath. When the lotion is cool, add .5 to 1 mL of inexpensive perfume.

### Part B: Formula 2

● Since it is difficult to dispense lanolin, it will be easier for students to weigh the beaker, then add anhydrous lanolin and re-weigh the beaker.

1. In a 300-mL beaker, mix the following substances.
   60 g of anhydrous lanolin (white)
   80 g of petrolatum
   60 g of glycerol
2. Heat the mixture to 70°C in a steam bath.
3. Remove the beaker from the bath and add 2 g of boric acid.
4. Add .5 to 1 mL of inexpensive perfume. Allow the lotion to cool.
5. Compare both lotions to a commercial hand lotion.

**Figure 38-1**



## QUESTIONS

**1.** What is an emulsion?

*A colloidal system consisting of one liquid dispersed in another liquid.*

**2.** What is a humectant?

*A moistening agent*

**3.** How does your hand lotion compare to that available at the drugstore?

*If the experiment is done properly, the hand lotion should be very similar to the drugstore variety.*

● Have a commercial variety of hand lotion available for comparison.

# Preparing Cleansing Cream

Cleansing cream and cold cream are oil and wax emulsions loosely holding a relatively large amount of water. When applied to the skin, the water-in-oil emulsion breaks down and water separates. As the cream is removed from the skin, dissolved particles of dirt, grease, make-up, and so on are removed with it. Also, as the water separates and evaporates, a cooling effect is produced on the skin.

The cleansing cream which you will make in this experiment is very similar in composition and method of preparation to commercial products. You can take it home and use it as you would any other cleansing or cold cream.

## PROBLEM

To prepare cleansing cream, a stable oil-water emulsion.

### Materials

Paraffin wax
Mineral oil
Borax (sodium borate)
Stearic acid
Perfume (inexpensive)
Stirring rod
Beakers
Thermometer
Bunsen burner
Small, wide mouth jar (empty cold cream jar, mustard jar, baby food jar)
Ringstand

## PROCEDURE

In this experiment you will prepare two solutions and mix them to form the final product. It will be necessary to mix them at exactly the right temperature.

1. Fill an 800-mL beaker ¾ full of water and heat it on a ringstand.
2. To a 400-mL beaker add 20 g of paraffin. Cut or break the paraffin into fine pieces.
3. Place the 400-mL beaker inside the 800-mL beaker. This hot water bath prevents heating the paraffin directly.

● A water bath is commonly used when it is not desirable to heat a substance directly.

4. Continue heating the water bath until all of the paraffin has melted. Adjust the burner flame so that the temperature of the paraffin does not go above 70°C.
5. Obtain a 1.5 g sample of stearic acid. Put it aside.
6. Add 125 mL of mineral oil to the beaker containing the melted paraffin.
7. Add the stearic acid sample to the beaker.

● Be careful not to heat the water to boiling as it may spatter over the top of the beaker.

8. Continue heating the water bath until the temperature of the wax-oil-stearic acid solution is again between 60 and 70°C. Maintain the solution at this temperature.
9. In a separate beaker heat 70 mL of water to 70°C.
10. Add 1.5 g of sodium borate (borax) to the water.
11. Stir until the borax is in solution.

12. Remove the beaker from the heat and allow it to cool to 60°C.
13. While stirring constantly, slowly add the water-borax solution to the first solution. Also add 3 to 5 mL of inexpensive perfume.
14. Allow the emulsion to cool while continuing to stir. Before it begins to harden pour it into the small jar.

● If the emulsion is not stirred, it will form a water layer on the bottom of the jar.

## QUESTIONS

1. Describe the appearance of your product.

*It should be similar to commercial cold cream.*

2. Compare your ingredients with those in commercial cold creams.

● Have a commercial variety of cleansing cream available for comparison.

● This is a fun experiment and students enjoy it. However, since you are preparing a fuel, great care should be taken to see that students are not careless.

**EXPERIMENT 40** Chapter 16

# Making Canned Heat: A Practical Gel

Rather than forming a true solution, particles in the range of 10 to 1000 A form a colloidal dispersion. Colloidal dispersions are usually translucent or opaque. One type of colloidal system consists of a solid dispersed in a liquid. When this type of system is fluid, as in paints, the dispersion is called a sol. When the motion is more restricted and the dispersion has a firmer consistency, it is called a gel. In this experiment you will make canned heat, a gel. Similar gels are used for cooking at campsites.

## PROBLEM

To prepare canned heat and test some of its properties.

### Materials

Denatured alcohol (ethanol)
Saturated calcium acetate solution
Beaker
Wire gauze
Glass stirring rod or wooden splints
Small can (tuna fish)
Graduated cylinder

**CAUTION** Extreme care should be observed at all times during this experiment.

## PROCEDURE

1. Place 15 mL of denatured alcohol in a 100-mL beaker.
2. While swirling the beaker, add 2 mL of saturated calcium acetate solution. Add the calcium acetate all at once, not drop-by-drop.
3. As soon as the calcium acetate has been added, stop swirling the beaker and set it aside.
4. After about 1 minute stir the contents of the beaker. Record your observations.

*The two solutions form a gel.*

5. Remove a small piece of the gel (about the size of a dime). Place it on a wire gauze and ignite it. Describe the results.

*The gel will burn with intense heat.*

6. Place a small amount of gel in a tuna fish can and ignite it. Describe the results.

*A confined heat source is produced.*

## QUESTIONS

**1.** Describe the gel.

*The gel should be clear and colorless.*

**2.** What happened when the gel was ignited?

*It burned with an intense heat and no smoke.*

**3.** What practical applications does a canned gel have?

*It can provide heat for cooking and even warmth under certain conditions.*

● There are many ways to prepare nylon. This is perhaps the easiest and most interesting.

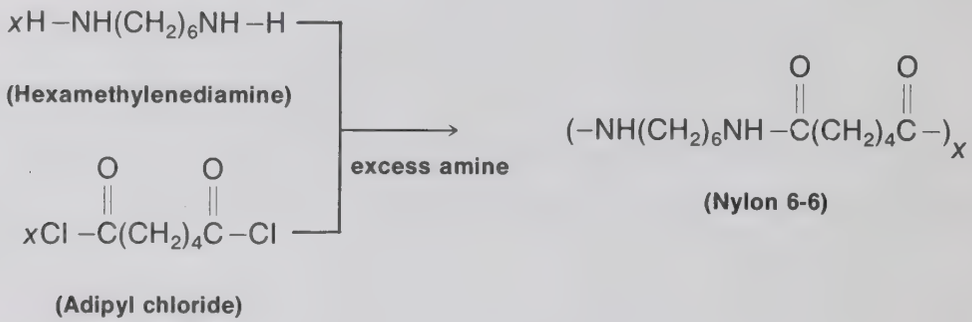
**EXPERIMENT 41**



# Preparation of Nylon

Nylon is a very long chain of repeating units of two basic compounds—an acid chloride and an amide. A compound composed of such repeating units is called a polymer. In this experiment you will prepare nylon.

The acid chloride used in this experiment is adipyl chloride. The amide will be provided by hexamethylene diamine. Since both members of the polymer have 6 carbons each, the nylon produced is referred to as "6-6" nylon. If sebacoyl chloride, a 10-carbon molecule, is used, nylon "6-10" is produced. The reaction is as follows:




## PROBLEM

To make nylon.

### Materials

● The adipyl chloride solution should be 0.25 *M* in hexane.

● The hexamethylene diamine solution should be 0.50 *M* in 0.50 *M* NaOH.

Adipyl chloride solution
Hexamethylene diamine solution
Acetone (in alcohol)
Paper clip
Beakers
Graduated cylinder

## PROCEDURE

1. Bend apart a paper clip, leaving a small hook at one end.
2. Pour 5 mL of hexamethylene diamine solution into a small beaker.
3. Slowly add 5 mL of adipyl chloride solution to the beaker. What do you observe?

*A white film immediately forms at the interface of the two solutions.*

4. Insert the paper clip into the beaker, hooking the film. Slowly lift the clip from the beaker. Record your observations.

*The polymer (nylon) can be pulled from solution like a string.*

5. Continue removing the nylon thread from the beaker until no nylon remains.
6. Place the nylon in a small beaker and wash it. Rinse the nylon with acetone. Allow it to dry. Examine the nylon which you have formed.

● Collect all excess chemicals, especially acetone. Do not permit students to dispose of these by pouring them down the sink. They are flammable.

## QUESTIONS

1. How does your polymer compare with a natural material like silk, for example?

*Answers will vary.*

2. Find out how other synthetic polymers are made. (Use the library.)

# The Chemical Weather Predictor

Certain chemical compounds are hygroscopic. This means that they readily absorb moisture from the air. This is commonly observed with table salt. Moisture causes the salt grains to adhere to each other. To prevent this, a few grains of rice can be added to the salt shaker. The rice absorbs the moisture from the air, keeping the salt dry.

Most hygroscopic substances change color when they absorb moisture. The dry, or anhydrous, form is a different color from the hydrate. In this experiment you will make a simple weather predictor using a hygroscopic compound.

## PROBLEM

To illustrate the hygroscopic nature of cobalt chloride.

● $CoCl_2$ solution—13 g/L
NaCl solution—6 g/L

### Materials

0.1 *M* cobalt chloride
0.1 *M* NaCl
Filter paper
Beaker

## PROCEDURE

1. Mix equal volumes of cobalt chloride solution and sodium chloride solution in a 250-mL beaker.
2. Dip a piece of filter paper into the solution, wetting it thoroughly. What color is the paper?

*The wet paper will be pink.*

3. Place the filter paper in a place where it will dry.
4. When the paper is dry, notice that it has changed color. What is the color of the anhydrous form?

*The paper is blue when dried.*

5. Place the paper near an open window, or on a wall outside of the classroom. You may wish to cut the paper into an appropriate shape—umbrella, cloud, and so on.
6. Watch the weather indicator. The slightest increase in atmospheric moisture will cause the paper to change color.

## QUESTIONS

**1.** What is the reason for adding NaCl to the solution?

*NaCl will help pull moisture more rapidly from the air than cobalt chloride alone.*

**2.** What color is the dry cobalt chloride?

*Blue.*

**3.** What color change is evident with the presence of moisture?

*It changes from blue to pink.*

**4.** What practical application could such a reaction have?

*Filter in gasoline lines, novelties, and so on*

● This is an excellent way to introduce hydration. You may wish to follow this experiment with a more quantitative study of dehydration of $CuSO_4 \cdot 5H_2O$.

## EXPERIMENT 43



# Making Raisins

The word raisin comes from the French *raisin sec* which means *dried grape*. During the process of making raisins, grapes are dehydrated. Most of the raisins produced in the United States are made from the Thompson seedless grape. There are basically four processes for making raisins.

**1.** Natural Raisins
These are dried in the sun without the addition of chemicals. They are dark brown.

**2.** Golden Bleached Raisins
To make these raisins, the grapes are first cured in a 0.5% NaOH solution. The sodium hydroxide serves two purposes. First, it removes the wax from the grape. Secondly, it cracks the skin to speed dehydration. The grapes are then exposed to fumes of burning sulfur. Then they are dried with artificial heat. These raisins are golden yellow.

**3.** Sulfur Bleached Raisins
These are processed in a manner similiar to those described in 2. However, they are dried in the sun. They are light beige to orange in color.

**4.** Lexia Raisins
To make Lexia raisins, the grapes are dipped in olive oil containing NaOH solution. The olive oil makes the skin tender. The NaOH acts as in 2. These grapes are also dried in the sun. They are medium to dark brown.

● Encourage students to perform this experiment at home. But warn them that they must not lose any of the grapes!

## PROBLEM

To investigate the process of dehydration by determining how much water is lost when raisins are made from grapes.

### Materials

Fresh, ripe, seedless grapes
Paper plate
Piece of cheesecloth large enough to cover the paper plate
Baking tray
Oven (or school incubator, radiator, heat lamp)
Beakers
Bunsen burner

## PROCEDURE

### Part A: Raisins in the Sun

1. Wash the grapes by dipping them into a large beaker of water. Blot the grapes dry with soft paper toweling. Be careful not to bruise them.
2. Select enough grapes to almost cover the bottom of the paper plate.
3. Determine the mass of the grapes and record this data in Table 43-1.
4. Spread the grapes evenly on the paper plate. Cover the plate with cheesecloth. Staple the edges of the plate to the cheesecloth.

5. Select a location outside the classroom, or at home, where the grapes can be exposed to direct sunlight. They should be placed so that air can freely circulate around them.
6. After four days, test the grapes by squeezing one lightly between your fingers. If there is no moisture left on your fingers, and if the grape springs back when released, they have dried enough. They should be soft and leathery.
7. Determine the mass of the grapes again. Be careful not to lose any if they have to be returned to the laboratory for this purpose. Record your data in Table 43-1.

**TABLE 43-1 • Moisture in Raisins**

| | |
|---|---|
| Mass of Grapes (Part A) | |
| Mass of Raisins | |
| Mass of Water Lost | |
| Mass of Grapes (Part B) | |
| Mass of Raisins | |
| Mass of Water Lost | |

### Part B: Making Raisins with Artificial Heat

1. Wash, dry, and determine the mass of the grapes as you did in Part A. Record the mass in Table 43-1.
2. Preheat the oven to about 150°F.
3. Dip the grapes into a large beaker of boiling water. Record what happens. Why is this step necessary?

*This breaks the skin of the grapes, making oven drying easier.*

4. Place the grapes in a single layer in the baking pan and put them in the oven.
5. Turn the grapes at least twice while they are being dried.
6. Dry the grapes until they are soft and leathery. This will probably require between 2 and 4 hours.
7. When the grapes are dry, remove them from the oven. After they have cooled, determine their mass. Record your data in Table 43-1.

## QUESTIONS

1. Calculate the mass of water lost in each process.

*Answers will vary.*

● The mass of water lost will vary. Those dried slowly will lose less water and have a leathery appearance.

● Correlation on this experiment should be very good.

2. Determine the percentage of water in a single grape (average).

3. Compare the physical characteristics of a raisin with a grape.

4. Devise your own procedure for making prunes.

5. Some chemical compounds, called hydrates, have water molecules chemically bound into their structure. One such compound is $CuSO_4 \cdot 5H_2O$. Devise an experiment to determine the percentage of water of hydration in this compound.

# Osmosis in Eggs

All plant and animal life is dependent upon materials in solutions. There must be some way for these solutions to pass through living tissue. The membranes of some tissues permit the passage of only certain small particles and liquids. These are called semipermeable membranes. The passage of water and other liquids through a semipermeable membrane tends to make the concentration of all substances equal on both sides of the membrane. This process is called osmosis. In osmosis, the tendency is to dilute the more concentrated side of a semipermeable membrane.

● Osmosis is usually applied to the movement of water across a membrane. Diffusion applies to the movement of dissolved materials across a membrane.

In this experiment you will demonstrate osmosis by using eggs. The egg membrane is a semipermeable membrane. To expose the membrane, you must partially dissolve the egg shell with vinegar. Vinegar is a weak solution of acetic acid in water. Since the concentration of materials inside the egg membrane is greater than in the water outside the membrane, water will enter the egg. This will cause the egg to increase in size. You will reverse the process by placing the egg in a more concentrated solution.

## PROBLEM

To demonstrate osmosis by using the membrane of an egg.

● Notice that this experiment requires several days—plan ahead!

### Materials

Raw eggs
Clear vinegar
Clear syrup
Large jars (wide mouth)

● Be sure that the jars are several times larger than the eggs; otherwise the eggs will swell and become lodged in the jars.

## PROCEDURE

1. Place an egg in a wide mouth jar. The egg should be much smaller than the mouth of the jar.
2. Add enough vinegar to more than cover the egg.
3. Place a cap on the jar to retain the vinegar odor.
4. Set the jar aside for two days. However, you should make observations frequently for the first 2 to 3 hours.
5. After two days, describe the effect of osmosis on the egg.

*The egg will swell to twice its original size.*

6. Carefully pour off the vinegar. Wash the egg several times by adding water to the jar and pouring it off. Remember, a thin membrane is holding the egg together—don't break it!
7. Add enough clear syrup to more than cover the egg.
8. Set the jar aside for two days.

9. After the second day, observe the effect of reverse osmosis. Describe the egg now.

*The egg will shrink.*

## QUESTIONS

1. What noticeable effect does the vinegar have on the egg shell within 1 to 2 hours?

*It dissolves the calcium in the shell.*

2. Describe the egg after treatment with vinegar.

*The egg swelled.*

3. What has passed through the egg membrane to cause the effect you observed?

*Water*

4. Describe the egg after treatment with clear syrup?

*The egg shrank.*

5. Explain your answer to question 4.

*The concentration is greater outside the egg membrane, thus water leaves the egg.*

**EXPERIMENT 45** Chapter 19 Exc. 12

# Detecting Calcium Propionate In Bread

One of the greatest problems in the food industry is spoilage. Bread is one of the foods that spoils very easily. Bread can be kept in a refrigerator to retard the growth of mold. Bread can also be made resistant to mold. Calcium propionate is the chemical most often used for this purpose. It is mixed with the dough before the bread is baked.

The bread used in this experiment must contain calcium propionate as a preservative. You could check for the presence of calcium propionate in bread by identifying the calcium ion, $Ca^{+2}$. However, there are other compounds in bread which also contain calcium. Therefore, you will identify the propionate group, since there are no other compounds in bread containing propionate. Your task is complicated by the fact that the amount of propionate in bread is very small. You will first convert the propionate into propionic acid. Then you will convert the propionic acid into an ester. An ester is an organic compound that is characterized by a pleasant odor. This particular ester has the odor of pineapple. You will use the pineapple odor as a basis for detecting the presence of propionate in bread.

● We will add sulfuric acid to precipitate the $Ca^{+2}$ and also to form propionic acid. Then, we will add ethyl alcohol that will react with propionic acid to form the ester, ethyl propionate.

## PROBLEM

To identify calcium propionate in bread.

### Materials

Bread
6 *M* $H_2SO_4$
95% ethyl alcohol
Beaker
Test tube
Bunsen burner
Stirring rod

● The reaction is:
$(C_2H_5COO)_2Ca + H_2SO_4 \rightarrow CaSO_4(s) + 2C_2H_5COOH$
$C_2H_5COOH + C_2H_5OH \rightleftharpoons H_2O + C_2H_5COOC_2H_5$
The equilibrium involved in esterification may not go very far to the right.

## PROCEDURE

1. Break up a slice of fresh bread and put it in a 250-mL beaker.
2. Add 100 mL of distilled water to the beaker.
3. Stir the mixture thoroughly with a stirring rod. Calcium propionate will dissolve in the water.
4. Pour off about 10 mL of the water solution into a test tube, leaving the solid bread behind.
5. Add 6 *M* $H_2SO_4$ to the tube, 1 mL at a time. Add no more than 5 mL of acid. What do you observe?

*There will be a cloudy precipitate of $CaSO_4$.*

● Be sure to use fresh bread containing calcium propionate.

**CAUTION** Several cautions should be noted in this experiment.
1. Advise the students to be careful handling the sulfuric acid.
2. Alcohol and ethyl propionate are flammable. Be careful when heating them. Be sure the bottles of alcohol are closed and out of the way.
3. Demonstrate the proper method of odor detection. Warn the students not to hold the tube beneath their noses.

6. Add 4 mL of ethyl alcohol to the tube and heat the tube gently and evenly.
7. Just before the liquid in the tube begins to boil, carefully wave some of the vapor toward your nose. Describe the odor of the vapor.

*Pineapple*

## QUESTIONS

1. Is there any evidence that calcium propionate is harmful when used as a preservative in bread?

*No*

2. What other preservatives can you name? Are any of these considered harmful?

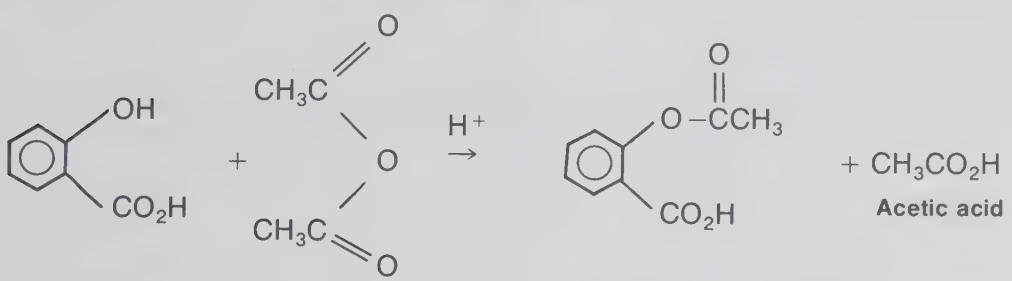
*Answers will vary.*

# Making Aspirin

Aspirin is one of the oldest and most useful drugs known. It is effective in reducing pain (analgesic) and it is one of the few drugs that effectively reduces fever (antipyretic). Aspirin is also one of the least expensive drugs, mainly because of the tremendously large volume of aspirin tablets produced each year—over 30 billion!

Aspirin is easily prepared by esterification. In this process an acid reacts with an alcohol, forming a compound called an ester. Esters are most readily identifiable by their odors. The odor of fingernail polish remover is caused by an ester. Wintergreen is another ester. Most fruit flavors are produced by esters. Aspirin is a special ester. It is prepared by reacting acetic acid with the alcohol group on salicylic acid. To make the reaction proceed more quickly, you will use the anhydride form of acetic acid. Your product will be acetylsalicylic acid, or simply aspirin. The reaction is as follows.

● Students should realize that this is a reaction between the acid group of the acetic anhydride and the -OH group of salicylic acid.




**Salicylic acid** **Acetic anhydride** **Acetylsalicylic acid** **Acetic acid**

## PROBLEM

To make aspirin.

### Materials

| | |
|---|---|
| Salicylic acid | Beaker |
| Acetic anhydride | Filter paper |
| 1% $FeCl_3$ | Graduated cylinders |
| $H_2SO_4$(concentrated) | Test tubes |
| Ice | Droppers |
| Thermometer | Bunsen burner |
| Erlenmeyer flask | Ringstand |
| Funnel | Trough |

## PROCEDURE

1. Determine the mass of approximately 3 g of salicylic acid as accurately as possible. Record the mass. Place the salicylic acid in a 250-mL Erlenmeyer flask.
2. Carefully and slowly add 6 mL of acetic anhydride to the flask.
3. Carefully add 5 drops of $H_2SO_4$ to the flask.

4. Swirl the flask to mix the reagents. Heat the flask in a water bath by placing it inside a larger beaker of water and heating to 80 to 90°C for at least 10 minutes.
5. Remove the flask from the water bath and allow it to cool to room temperature.
6. Place the flask in a trough of ice and add 40 mL of water to the flask. Record your observations.

*A white crystalline solid forms.*

If the filtrate is brown, the students must wash the aspirin and filter again.

7. Filter the crystals, discarding the filtrate. Wash the crystals twice with small quantities of cold water. These crystals are your aspirin.
8. Blot the crystals between paper towels. Then let them dry in the open. Recover as much of the aspirin as possible and determine the mass.
9. To test for any unreacted salicylic acid, add a few crystals of salicylic acid to a test tube. Dissolve the sample in 5 mL of water.
10. Add 1 drop of $FeCl_3$ solution. Note the color produced by the salicylic acid.

*Pink, violet, or green*

11. Repeat, using a few crystals of your aspirin instead of the salicylic acid. Does the aspirin contain any unreacted salicylic acid or is it pure?

*Answers will vary.*

The students were asked to determine the mass of the sample of salicylic acid used and the mass of the product in order to permit them to calculate a percent yield of product. If you elect to have them do this, they should perform the following calculations.
Determine the number of moles of salicylic acid used (mol. mass of salicylic acid = 138). Determine the theoretical number of moles of aspirin produced.
Determine the theoretical number of grams of aspirin produced (mol. mass of aspirin = 180).

$$\% \text{ yield} = \frac{\text{g aspirin (obtained)}}{\text{g aspirin (theoretical)}} \times 100$$

## QUESTIONS

1. Describe the product which you have formed.

*White, crystalline solid*

2. Although you shouldn't taste anything in the laboratory without your teacher's direction, touch a finger to your aspirin and then to your tongue. How does it taste? How does aspirin taste?

*Bitter*

EXPERIMENT 47

Chapter 17

# Acids and Bases

● This experiment introduces students to acid-base titration using an indicator. You will need to demonstrate the proper use of the buret and the technique of titration.

Acids react with bases and, in the presence of a suitable indicator, a color change is evident when neutralization occurs. This provides an excellent way to express the strength of a basic solution—how much of it is required to neutralize an acid used as a reference solution. In this experiment you will use vinegar as your reference solution. You will use phenolphthalein as the indicator. Using this reference acid, you will determine the relative strength of two substances, sodium bicarbonate (baking soda) and sodium tetraborate (borax). Although these compounds are not bases (they are actually salts), they act as bases in aqueous solution.

## PROBLEM

To determine the relative strengths of sodium bicarbonate and sodium tetraborate by titrating them with vinegar.

### Materials

Vinegar
5% sodium bicarbonate solution
5% sodium tetraborate solution
Phenolphthalein
Buret
Erlenmeyer flasks
Graduated cylinder

## PROCEDURE

1. Fill a clean buret with vinegar.
2. Place 15 mL of sodium bicarbonate solution in a small Erlenmeyer flask.
3. Add 2 or 3 drops of phenolphthalein indicator solution to the sodium bicarbonate. What color is the solution?

   *Phenolphthalein is pink in basic solution.*

4. While swirling the flask, carefully add vinegar a drop at a time until the solution becomes colorless. (See Fig. 47-1) Record the volume of vinegar used.

   Volume of vinegar. *Answer will vary.*

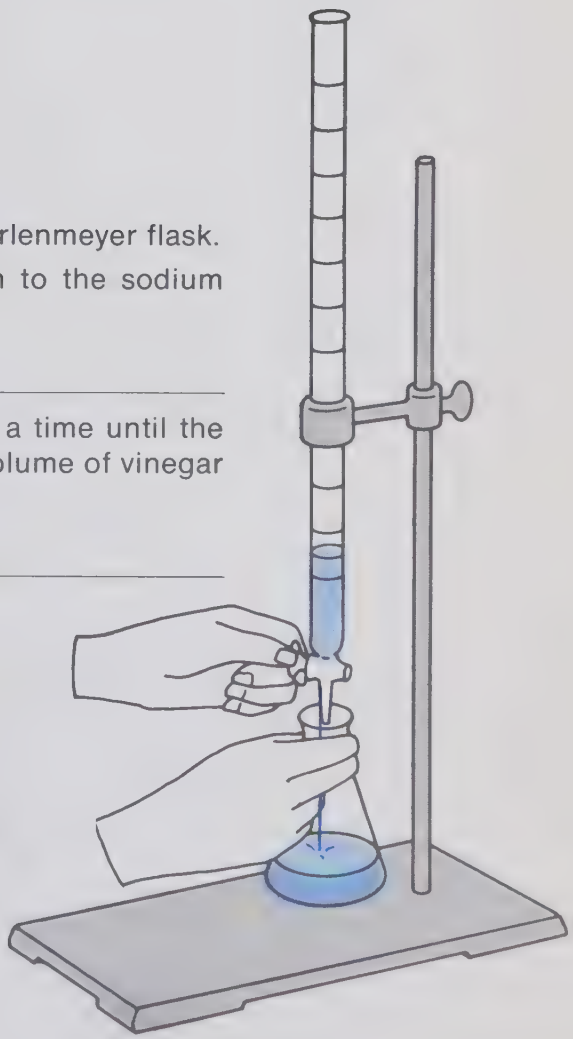


**Figure 47-1**

● You can have students repeat the experiment using a 5% solution of household ammonia.

5. Repeat the experiment, using sodium tetraborate (borax) solution. The buret should be refilled with vinegar before you titrate. Record the volume of vinegar used.

Volume of vinegar. *Answer will vary.*

## QUESTIONS

1. What volume of vinegar is required to titrate the sodium bicarbonate solution? What volume of vinegar is required to titrate the sodium tetraborate solution?

*Answers will vary.*

2. What does the experiment tell you about the relative strengths of the two basic solutions?

*One "base" requires more acid than the other for neutralization, hence it is a stronger base.*

3. Which solution is more basic? How many times more basic?

*Sodium tetraborate is about six times more basic.*

## EXPERIMENT 48

Chapter 17

# Natural Indicators

● In Experiment 47 the students did a simple titration. In this experiment they examine some natural indicators. This experiment is designed to remove the mystery from the indicator—it is simply a dye that changes color at different pH's.

One of the characteristic properties of an acid is that it will react with a base to produce a salt. In Experiment 47 you used an indicator to determine when neutralization of a base by an acid was complete. In this experiment you will use several indicators that occur naturally.

## PROBLEM

To prepare a natural indicator and use the indicator to estimate the strength of an acid by titration.

### Materials

Red cabbage
Vinegar
Baking soda (sodium bicarbonate)
Beakers
Test tubes
Buret
Universal pH paper
Graduated cylinder
Erlenmeyer flask
Strainer
Bunsen burner
Ringstand

● Be sure the cabbage is *red*.

● Other natural indicators:
blueberry juice
beet juice
carrot juice
grape juice
skins of elderberries
skins of grapes
purple hollyhock (flower)
blue iris (flower)
purple dahlia (flower)

## PROCEDURE

### Part A: Preparing the Red Cabbage Juice Indicator

● You might wish to boil ½ cabbage in a large beaker and provide indicator for the class.

1. Cut a section (about ¼) of a small red cabbage. Chop the red cabbage into fine pieces.
2. Place the red cabbage in a large beaker containing 500 mL of water. Heat the mixture until it just begins to boil.
3. Remove the beaker from the heat and allow it to cool. Strain the juice into a second beaker. This is the indicator solution.

### Part B: Titration

1. In a beaker, add 5 g of sodium bicarbonate to 250 mL of distilled water. Use this as your stock solution.
2. Place 50 mL of the stock solution in a 250-mL Erlenmeyer flask.
3. Add about 15 mL of the red cabbage juice indicator to the flask. Measure the pH with universal pH paper. Record your data in Table 48-1.
4. Fill a buret with vinegar.
5. Add vinegar, a drop at a time, from the buret until a color change is evident. Record the color change in Table 48-1.
6. Measure the pH again and record it. Record the volume of vinegar required to cause the color change.

● Encourage students to devise their own method for preparing an extract to use as an indicator.

● A pH meter may be used in place of the pH paper.

7. Repeat the experiment, using several of the natural indicators supplied by your teacher. Record your results with each indicator.

**TABLE 48-1 • Indicators and Titration**

| INDICATOR | pH OF $NaHCO_3$ PLUS INDICATOR | COLOR CHANGE | pH AFTER COLOR CHANGE | VOLUME OF VINEGAR |
|---|---|---|---|---|
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |
| | | | | |

## QUESTIONS

1. At what pH does each indicator show a color change?

2. How does the volume of vinegar used in each case compare with the pH after the color change?

*The more vinegar used, the more acid the solution.*

# Testing Antacid Tablets

You are probably familiar with TV advertising for antacids. These products are used to neutralize stomach acids. How do antacids work? Most antacids contain carbonates or bicarbonates. These have a neutralizing effect on acids. They also produce $CO_2$ when they react with acids. The buildup of this gas in the stomach causes the familiar burp, which in itself provides some relief. Many antacids are insoluble or only slightly soluble in water. This reduces the possibility of the antacid being absorbed into the bloodstream. This would produce excess base in the blood, a condition called alkalosis. Calcium carbonate, $CaCO_3$, magnesium carbonate, $MgCO_3$, and magnesium oxide, MgO, are examples of insoluble compounds that might be found in antacids. In this experiment you will neutralize HCl with various antacids. You will also compare their relative effectiveness in neutralizing the acid.

## PROBLEM

To compare the effectiveness of several antacids in neutralizing acid.

### Materials

1 *M* HCl
Congo red indicator
Several brands of antacids
Buret
Erlenmeyer flask
Stirring rod

● Any commercial antacids in tablet, liquid, or powdered form can be used.

## PROCEDURE

In order to directly compare the various antacid tablets, you will use 1 g of each antacid and titrate it with 1 *M* HCl.

1. Fill a clean buret with 1 *M* HCl.
2. Select one of the antacid products. If it is in solid or tablet form, crush it to a powder and obtain a 1-g sample.
3. Place the sample in a 250-mL Erlenmeyer flask.
4. Add 15 mL of water and stir until some of the antacid is in solution. The solution will probably appear cloudy.
5. Add 4 to 5 drops of congo red indicator to the flask. Record the color of the solution.

*Orange*

● It will make the experiment more interesting to the class if you will call HCl "stomach acid." Actual stomach acid is HCl with a pH ranging from 0.9 to 1.5. Point out to the students that the acid used in this experiment is not as strong as typical stomach acid.

6. Carefully add HCl from the buret, drop-by-drop, until a color change is noted. Record the color of the solution.

*Blue*

● If students are not familiar with titration from previous experiments, it will be necessary to demonstrate this technique.

7. Record in Table 49-1, the volume of acid used. Since 1 mL of HCl has a mass of approximately 1 g, you can assume that the number of milliliters of acid used is equal to the number of grams of acid reacting with 1 g of the antacid.
8. Repeat the procedure for each antacid assigned to you.

**TABLE 49-1 • Testing Antacid Tablets**

| ANTACID | VOLUME OF ACID USED (mL) | MASS OF ACID USED (g) |
|---|---|---|
| | | |
| | | |
| | | |
| | | |
| | | |
| | | |

## QUESTIONS

1. What is the color of congo red in basic solution? What color is it in acid solution?

*Orange in basic and blue in acid*

2. Which antacid might be considered the best acid neutralizer? Which was the least effective?

*Answers will vary.*

**EXPERIMENT 50** Chapters 17, 19, 20

# Testing for Vitamin C in Foods

Vitamin C, also called ascorbic acid, is an important nutrient required by human beings. Its primary function is to activate certain enzymes which catalyze oxidation-reduction reactions in the body. It is also necessary for the repair of injured tissue. Vitamin C is found naturally in many foods, especially citrus fruits. It is also added as a preservative or nutrient supplement to many foods. Vitamin C reacts with iodine in solution. Iodine also reacts with starch to give a characteristic purple color. Therefore, you can use the starch test to test for vitamin C.

● We assume that the iodine will not react with the starch until it has reacted with *all* of the vitamin C. At that point the excess $I_2$ will react with the starch.

You will determine the amount of iodine required to react with vitamin C by doing a titration. You are not concerned with determining the exact amount of vitamin C in a sample of food. You are interested in comparing the vitamin C content of several types of foods, especially fruit juices.

## PROBLEM

To estimate the relative amounts of vitamin C in several products.

### Materials

● Don't use grape juice. It is the same color as the blue color in the starch test. The vitamin C tablets can be of a low concentration.

Fresh fruit juices
Canned fruit juices
Frozen fruit juices (reconstituted)
Multiple vitamin tablets
Vitamin C tablets
Canned beverages
Powdered beverage preparations (reconstituted)
1 *M* HCl
Iodine solution
Starch solution
Stirring rod
Buret
Graduated cylinder
Erlenmeyer flask
Mortar and pestle

● Prepare the iodine solution by dissolving 0.6 g of $I_2$ in 100 mL of ethyl alcohol. Add 0.6 g of KI and dilute the solution to 1 liter with water. The starch solution is 1%. First make a paste with the starch in 25 to 30 mL of boiling water. Then add hot water to bring the volume to 1 liter.

## PROCEDURE

Follow step 1 or step 2, depending on the nature of your sample.

**1.** a. If your sample is a solid, determine the mass and record the mass in Table 50-1.
b. Mash or grind the sample with a mortar and pestle.
c. Dissolve the sample in about 30 mL of water and transfer it to a 250-mL Erlenmeyer flask.

**2.** If you use a liquid sample, measure a 25- to 30-mL sample into a 250-mL Erlenmeyer flask. Record the volume in Table 50-1.

**3.** Add 2 drops of HCl to the sample and mix it thoroughly.

**4.** Add 1 mL of starch solution and mix by swirling the flask.

5. Using a buret, slowly add the iodine solution to the flask until a dark color appears. When the dark color remains for 10 to 15 seconds, you have reached the end-point. Record the volume of iodine solution used.

6. Repeat the procedure using the samples suggested by your teacher.

● Stress to the students that this is merely *comparing* vitamin C content of various foods.

**TABLE 50-1 • Testing for Vitamin C**

| SAMPLE | MASS OR VOLUME USED | VOLUME OF IODINE USED |
|---|---|---|
| | | |
| | | |
| | | |
| | | |
| | | |
| | | |

## QUESTIONS

1. Compare the vitamin C content of your samples. To do this, divide the volume of iodine used by the mass (or volume) of the sample.

_____

_____

_____

_____

● Students can place their results on a classroom chart or on the chalkboard.

2. Compare your results with those of other students in the class.

_____

_____

_____

_____

● If you are interested in doing a more quantitative experiment, any college-level laboratory manual in chemistry will suggest procedures.

## EXPERIMENT 51

Chapter 17

# Etching with Hydrofluoric Acid

● This experiment works very well but it must be done with teacher supervision. You should use liquid HF only as an absolute last resort. Under no circumstances should students be allowed to use HF at their desks.

The strength of an acid refers to the concentration of hydrogen ions that the acid produces in a solution at equilibrium. A strong acid releases a large number of hydrogen ions. A weak acid releases very few. However, a weak acid can be very corrosive. To stress this point, you will investigate hydrofluoric acid, HF. Chemically, hydrofluoric acid is a weak acid. Yet it is so corrosive it must be stored in wax or plastic containers since it will dissolve a glass bottle.

In this experiment you will use hydrofluoric acid to etch glass. The major component of glass is calcium silicate, $CaSiO_3$. The reaction of glass with hydrofluoric acid follows.

$$CaSiO_3 + 6HF \rightleftharpoons CaF_2 + SiF_4(g) + 3H_2O$$

The silicon tetrafluoride ($SiF_4$) produced is a gas.

● The students will produce HF by reacting calcium fluoride with sulfuric acid.
$CaF_2 + H_2SO_4 \rightleftharpoons 2HF + CaSO_4$
An alternate method is to use liquid HF. However, the acid is very corrosive to the skin and can cause painful, slow-healing burns. For this reason, *liquid HF must be used only by the teacher.*

## PROBLEM

To produce hydrofluoric acid and use it to etch glass.

### Materials

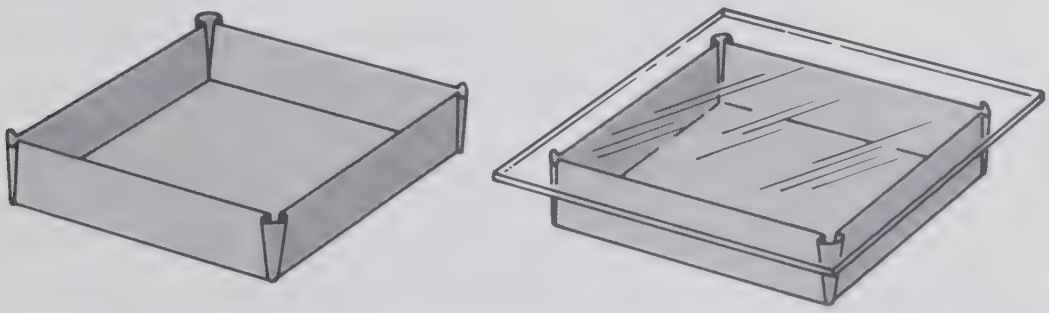
Calcium fluoride
$H_2SO_4$ (concentrated)
Square piece of lead foil (7.5 cm on a side)
Glass plate
Round toothpick
Tongs

## PROCEDURE

● *This experiment must be carried out in a fume hood.*

1. Form a small dish by turning up the edges of the square piece of lead foil. Place the glass plate on the top of the dish. The plate should completely cover the dish. (See Fig. 51-1)

**Figure 51-1**



● A teflon or heavy plastic container may be substituted for the lead foil. The glass plate should be slightly larger than the container.

2. Using tongs, dip the plate quickly into a large beaker of melted paraffin or wax. With a little practice you should be able to get a very thin coat of wax on the glass plate. Allow the wax to harden.
3. Using a toothpick, make a design by scratching through the wax. Be sure to scratch through the wax to the surface of the glass.

● Prepare both the beaker of melted wax and the beaker of hot water for the class to use. If the wax coating on the plate is too thick, have the student immerse the plate in the hot water to remove the wax and then repeat the procedure.

Supervise this!

If necessary, students can fan the container slightly with a Bunsen burner flame.

If you use liquid HF, place the HF under a hood in a large plastic trough. Carefully, using tongs, immerse the coated plates with the student's designs in the HF for about 5 minutes. Take the plates out with the tongs and immediately immerse the plate in the beaker of hot water. Avoid dripping the HF.

4. Place the lead dish under a fume hood. Add 2.0 g of calcium fluoride to the dish. Add enough concentrated $H_2SO_4$ to just cover the calcium fluoride.
5. Place the glass plate on the dish, design surface down, so that the design is exposed to the hydrofluoric acid vapors produced by the reaction. Do not breathe the vapors.
6. After 5 to 10 minutes, use the tongs to remove the glass plate. Rinse the plate in a large beaker of hot water to melt the wax.
7. Rub the remaining wax off with a paper towel.

## QUESTIONS

1. What is the appearance of the glass plate?

*The design should be etched into the glass.*

2. Why is hydrofluoric acid considered a weak acid?

*It produces few hydrogen ions in solution.*

EXPERIMENT 52

Chapter 18 Exc. 11

# Iron in a Razor Blade

In Experiment 47 you titrated an acid with a base to determine the relative strength of the base. Titration is a very important analytical technique. In this experiment you will do a titration involving an oxidation-reduction (redox) reaction. Iron, in the +2 oxidation state, will be titrated with an oxidizing agent to yield iron in the +3 oxidation state. Using a suitable indicator, the volume of oxidizing agent required for the reaction can be measured, and the number of moles of reactants easily determined. The dichromate ion, $Cr_2O_7^{-2}$, will serve as the oxidizing agent. The net ionic reaction is as follows.

$$Cr_2O_7^{-2} + 6Fe^{+2} + 14H^+ \rightleftharpoons 6Fe^{+3} + 7H_2O + 2Cr^{+3}$$

Notice that 6 moles of $Fe^{+2}$ are effectively oxidized by 1 mole of $Cr_2O_7^{-2}$.

## PROBLEM

To determine the percent of iron in a razor blade.

### Materials

Razor blade (used double-edged)
3 *M* $H_2SO_4$
85% $H_3PO_4$
Diphenylaminesulfonic acid indicator
0.016 *M* $K_2Cr_2O_7$
Alcohol
Dropper
Filter paper
Funnel
100-mL volumetric flask
Buret
250-mL Erlenmeyer flasks
Hotplate

● Diphenylaminesulfonic acid —add 2 g of the sodium salt and dilute to 1 liter.

● The 0.016 *M* $K_2Cr_2O_7$ is 4.7 g/L.

## PROCEDURE

### Part A: Preparing the Iron Sample

1. Carefully wash the razor blade with alcohol and blot it with paper toweling.
2. Determine the mass of the blade as accurately as possible. Record the mass in Table 52-1.
3. Place the blade in a 250-mL Erlenmeyer flask.
4. Add 30 mL of 3 *M* $H_2SO_4$ to the flask.
5. Place the flask under a hood and warm it gently until the blade dissolves. Hydrogen gas will be produced. If the reaction becomes very vigorous, remove the flask from the heat until the reaction subsides.
6. Place a piece of filter paper in a funnel. Place the stem of the funnel into a 100-mL volumetric flask.

7. Carefully filter the sample to remove any undissolved carbon. Rinse the Erlenmeyer flask with distilled water. Pour the rinse through the filter. Be careful not to add more than 100 mL of liquid to the volumetric flask.
8. Using a dropper, add distilled water to the volumetric flask until the volume is exactly 100 mL. This is your iron solution.
9. Clean a 50-mL buret. Rinse it with 10 mL of the iron solution.
10. Fill the buret with the iron solution and carefully measure 25 mL of the solution into each of two 250-mL Erlenmeyer flasks.
11. Add 50 mL of distilled water to each flask.
12. Add 5 mL of 85% phosphoric acid to each flask.
13. Add 6 drops of diphenylaminesulfonic acid indicator to each flask.
14. Stopper each flask and set them aside.

● Phosphoric acid is used to complex the $Fe^{+3}$ ion, which otherwise might interfere and produce a premature end point.

### Part B: Titrating the Sample

1. Rinse the buret several times with distilled water. Rinse it once with 10 mL of $K_2Cr_2O_7$ solution.
2. Fill the buret with the $K_2Cr_2O_7$ solution.
3. Carefully titrate one iron sample with $K_2Cr_2O_7$. The permanent appearance of a purple color is the end point. Record, in Table 52-1, the volume of $K_2Cr_2O_7$ used.
4. Repeat the titration with the second flask. Record the volume of $K_2Cr_2O_7$ used. The volume of $K_2Cr_2O_7$ used in each titration should be nearly the same. If not, prepare another sample and repeat the titration. Use the two volume readings which most closely agree to answer the questions.

**TABLE 52-1 • Iron in a Razor Blade**

| | |
|---|---|
| Mass of Razor Blade | |
| Volume of $K_2Cr_2O_7$ Used (Sample 1) | |
| Volume of $K_2Cr_2O_7$ Used (Sample 2) | |

## QUESTIONS

1. Determine the number of moles of $K_2Cr_2O_7$ needed to titrate each sample.

*Volume of $K_2Cr_2O_7$ (mL) x 0.016 moles/1000 mL*

**2.** Determine the number of moles of Fe present in each sample.

*Moles of $K_2Cr_2O_7$ x 6*

**3.** Determine the mass of Fe present in each sample.

*Moles of Fe x 58.8 g/mole*

**4.** Calculate the percent of Fe in the razor blade. Do a separate calculation for each sample.

*(Mass of Fe in sample x 4/mass of razor blade) x 100*

● The sample was diluted—students are using ¼ of the original sample.

# Benzoic Acid and Sodium Benzoate in Food

Benzoic acid and sodium benzoate are often found in catsup, jellies, and jams as preservatives. These compounds inhibit the growth of fungi and bacteria that cause spoilage. Benzoic acid can be separated and extracted from a sample of food by using chloroform. It is soluble in this solvent but only slightly soluble in water. Sodium benzoate, however, is soluble in water. In this experiment you will extract benzoic acid from food and compare it to a sample of pure benzoic acid. You will also attempt to reclaim sodium benzoate that you will have added to a sample of food.

● Caution students about the use of chloroform.

## PROBLEM

To separate benzoic acid and sodium benzoate from samples of food. To identify benzoic acid and sodium benzoate.

### Materials

● Obtain samples of food containing benzoic acid and similar samples without the acid or benzoate.

Sodium benzoate
Chlorofrom
Ammonia water
1.0% iron(III) chloride
Benzoic acid
Cheesecloth
Evaporating dish
Beakers
Food samples (catsup, jelly, jam)
Funnel
Filter paper
Test tube
Bunsen burner
Separatory funnel
Stirring rod

## PROCEDURE

● Ether may be used in place of chloroform. However, while handling ether, no open flames should be used.

Prepare two samples for this experiment. Catsup, jelly, or jam containing benzoic acid will be one sample. Benzoic acid will be listed as an ingredient on the container. The other sample will be the catsup, jelly, or jam without benzoic acid or sodium benzoate. To this sample add 2 to 3 g of sodium benzoate. Then, analyze both samples according to the following procedure.

### Part A: Separation and Extraction

1. Mix 15 mL (1 tablespoon) of your sample with 30 mL of water in a small beaker.
2. Stir until the sample is thoroughly mixed.
3. Strain the mixture through cheesecloth, collecting the liquid in another beaker.
4. Filter the liquid, collecting the filtrate in a beaker.
5. Place the filtrate in a separatory funnel and add 10 mL of chloroform. Shake the funnel to mix the liquids. Allow the liquids to separate into two layers.

● If ether is the solvent, it will be the top layer.

6. Siphon off the chloroform (the lower layer) into an evaporating dish. Allow the solvent to evaporate. This must be done under a hood.

7. When all of the solvent has evaporated, place the white, crystalline residue in a small test tube.

**Part B: Identification**

1. Compare the color, odor, and appearance of your sample with a pure sample of benzoic acid.

2. Add 4 to 5 drops of ammonia water to the test tube containing your sample and dissolve the residue.
3. Heat the tube *slowly* until it is dry. Do not allow the sample to spatter.
4. When the tube is cool, add 4 to 5 drops of water and 2 drops of iron(III) chloride solution.
5. Record any evidence of a reaction. If benzoic acid or sodium benzoate is present, a light brown precipitate of iron(III) benzoate will form.

*A precipitate should form.*

## QUESTIONS

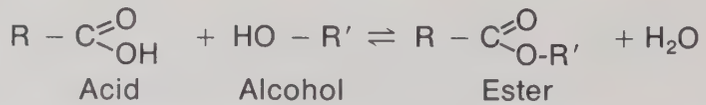
1. Describe the crystalline residue you obtained from each sample.

2. Describe the results of your tests for the identification of benzoic acid and sodium benzoate.

*The residues should appear similar to pure benzoic acid.*

# Artificial Flavoring

Many foods such as ice cream, candy, syrup, and others contain artificial flavors. Many of these flavoring additives are esters. An ester is an organic compound that is produced when an organic acid reacts with an alcohol. Esters are named by combining the names of the alcohol and the acid from which they are synthesized. For example, ethyl alcohol and benzoic acid produce the ester ethyl benzoate. Ethyl alcohol and decanoic acid produce ethyl decanoate. Methyl alcohol and salicylic acid produce methyl salicylate. The following general reaction shows the formation of an ester. R represents the carbon group characteristic of the acid. R′ represents the carbon group of the alcohol.




Esters have pleasant odors that help to identify them. Some esters occur naturally and others are produced artificially. The chart lists some common esters and their odors.

● The students can use the names of the esters given in the chart to deduce the names of the esters formed by reacting these alcohols and acids. They will need some help.

| ESTER | ODOR |
|---|---|
| Ethyl benzoate | Fruity |
| Ethyl butyrate | Pineapple |
| Ethyl cinnamate | Peach |
| Ethyl decanoate | Grape |
| Ethyl hexanoate | Pineapple/Banana |
| Ethyl isobutyrate | Apple |
| Isoamyl butyrate | Apricot |
| Isoamyl salicylate | Pineapple |
| Methyl salicylate | Wintergreen |
| Methyl butyrate | Apple/Banana |
| Methyl propionate | Rum |

## PROBLEM

To prepare several esters and note their characteristic odors.

### Materials

● Caution the students about using $H_2SO_4$.

Glacial acetic acid
Salicylic acid
Butyric acid
Decanoic acid
Amyl alcohol
n-Octyl alcohol
Methyl alcohol
Ethyl alcohol
Isoamyl alcohol
$H_2SO_4$ (concentrated)
Beaker
Test tubes
Hotplate
Aluminum foil
Graduated cylinders

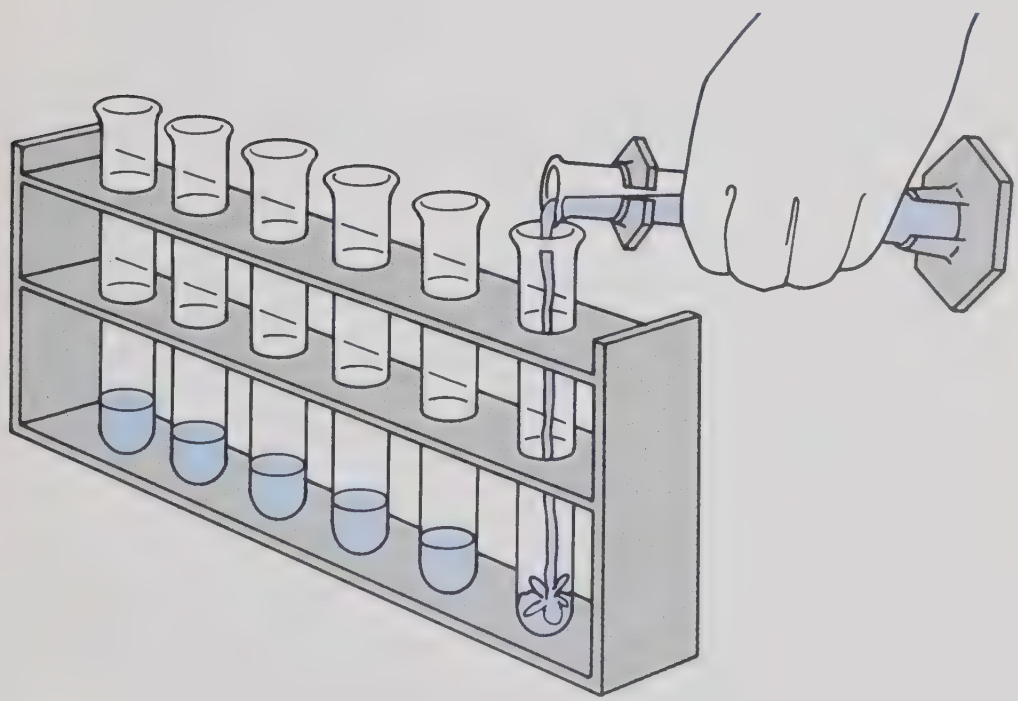
## PROCEDURE

1. Add 375 mL of water to a 500-mL beaker. Bring the water to a boil.
2. Place 6 test tubes in a test tube rack. Number the tubes 1, 2, 3, and so on.
3. Place 2 mL of each reagent into the appropriate tube, as indicated in the following chart. See Figure 54-1. (**Carefully** note the odor of each reagent as you add it.)

● Remind students of the proper way to smell reagents.

| TUBE | ALCOHOL | ACID |
|---|---|---|
| 1 | isoamyl | butyric |
| 2 | isoamyl | salicylic |
| 3 | amyl | glacial acetic |
| 4 | n-octyl | glacial acetic |
| 5 | methyl | salicylic |
| 6 | ethyl | decanoic |

**Figure 54-1**



4. Carefully add 1 mL of $H_2SO_4$ to each tube.
5. Stir each mixture slightly by tapping the bottom of the tube with your finger.
6. Place tubes 1 and 2 into the beaker of boiling water.
7. After a few minutes, remove the tubes from the water bath. Check each tube for the presence of an odor. Record your observations in Table 54-1.
8. If no odor is detected in a tube, cover the top with aluminum foil. Set the tube aside in a warm place for several hours. Check occasionally for the odor of the ester. Record your observations.
9. Repeat steps 6 through 8 for the remaining tubes.

● Show students this technique. Do not permit them to cover the top with a finger to shake the tube!

● Some of these esters have more distinct odors than others. Students should smell the reaction as soon as it begins, since some of the esters are "short-lived" and are gone in a few minutes.

**TABLE 54-1 • Esters**

| TEST TUBE | OBSERVATIONS |
|---|---|
| 1 | |
| 2 | |
| 3 | |
| 4 | |
| 5 | |
| 6 | |

## QUESTIONS

**1.** Describe the odor of the ester produced in each tube.

**2.** Write the reaction for preparing one of the esters.

**3.** Name the esters that you produced.

*Isoamyl butyrate; isoamyl salicylate; amyl acetate; n-octyl acetate; methyl salicylate; ethyl decanoate.*

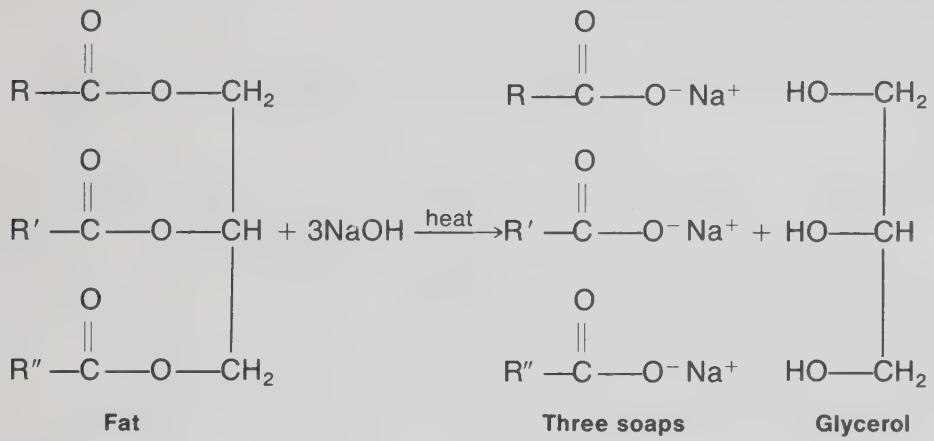
## EXPERIMENT 55

Chapter 20 Exc. 12

# Preparing Soap

● This experiment can be done with other ester experiments, since a fat is an ester of a long-chain fatty acid and glycerol.

Soap has been used for cleaning purposes for a long time. In the Middle Ages soap was prepared by boiling fats and oils in a water solution of the ashes from burnt wood or seaweed. Pioneer American families prepared their own soap by heating kitchen fat and lye (sodium hydroxide). Fats and sodium hydroxide will react when heated to form glycerol and soaps. This reaction is shown below.




(R, R′, and R″ represent long-chain hydrocarbons.)

This process is called saponification. The soaps produced are metallic salts of long-chain fatty acids.

## PROBLEM

To produce soap by saponification.

### Materials

Fat, oil, or lard
6 *M* NaOH
Sodium chloride
Ethyl alcohol
Beakers
Bunsen burner
Ringstand
Stirring rod
Test tube

● Lard works well and is inexpensive.

● **CAUTION** Ethyl alcohol is flammable.

## PROCEDURE

1. Place 100 mL of water in a beaker and bring the water to a boil. Lower the flame and add 25 g of sodium chloride. Stir the solution. Keep it warm over the low flame. You will need it later in the experiment.
2. Place 10 g of fat, lard, or oil in a 250-mL beaker.
3. Add 15 mL of 6 *M* NaOH to the beaker containing the fat. Also add 60 mL of ethyl alcohol.
4. Place the beaker in a water bath of boiling water.
5. Stir the contents of the beaker. If some of the volume is lost by evaporation, add distilled water to maintain the original volume.

● Remind students how to prepare a water bath with a large beaker.

6. Heat the solution for at least 30 minutes. After that time, remove a small amount of the solution from the beaker with a wooden splint or a stirring rod. Dip the splint or rod into a test tube ½ full of warm water. Shake the tube thoroughly. Record your observations.

*Suds should form.*

7. If suds form and no fat globules are present, add 20 mL of water to the beaker. Then add 50 mL of the warm salt solution you prepared in step 1. (If no suds form or fat globules are present continue to heat the solution. After 15 minutes, repeat step 7.)
8. Remove the beaker from the bath and let it cool. Record your observations.

*The soap will float to the top of the beaker and form a cake.*

● If the soap is granular, filter it and blot it with paper towels.

9. Lift out the soap formed in the reaction.

## QUESTIONS

1. Describe the soap you produced.

*The soap should be clear and white.*

● Do not permit students to wash their skin with this soap.

2. Your soap should not be used on the skin because it may contain some unreacted alkali (NaOH). How could you test for the presence of unreacted alkali?

*Dissolve a small bit of the soap in water and test the soap solution with litmus paper.*

3. Compare your soap, in appearance only, to a commercially prepared soap.

*It will be similar.*

EXPERIMENT 56 Chapter 20 Exc. 12

# Preparing Dreft: A Detergent

Unlike soaps, detergents are not affected by the metal ions in hard water. Thus, they will form suds in the presence of calcium and magnesium ions. In this experiment you will prepare a detergent, and test its properties. The detergent is sodium lauryl sulfate and has the following formula.

● This is the big difference between a soap and a detergent: the detergent will form suds in hard water, the soap will not.

$$CH_3\,(CH_2)_{10}CH_2 - OSO_3^- Na^+$$

## PROBLEM

To prepare a synthetic detergent and test its properties.

### Materials

$H_2SO_4$ (concentrated)
Dodecanol (lauryl alcohol)
6 *M* NaOH
Phenolphthalein
Beakers
Graduated cylinder
Stirring rod
Funnel
Filter paper
Calcium chloride
Test tubes

● Caution students about the use of $H_2SO_4$.

## PROCEDURE

1. Obtain two 250-mL beakers and label them beaker A and beaker B.
2. Place 15 g of dodecanol in beaker A.
3. Slowly add 10 mL of $H_2SO_4$ to beaker A. Use caution!
4. To beaker B add 30 mL of NaOH and 3 drops of phenolphthalein. The solution should be colorless.
5. With constant stirring, add the acid solution in beaker A to the basic solution in beaker B. Describe the reaction.

*The mixture becomes red and finally decolorizes. A precipitate forms.*

6. Filter the solution and allow the precipitate to dry on the filter paper. This is your detergent.
7. When the detergent is dry, add a small amount to a test tube ½ full of water. Shake the tube and test for suds formation. Record your observations.

*Suds should form.*

8. Add a small amount of the detergent to a test tube ½ full of water that contains a pinch of calcium chloride. Shake and record your observations.

*Suds should form.*

## QUESTIONS

1. How does the detergent differ from soap in its action?

*It forms suds in hard water.*

2. What advantages does a detergent have over a soap?

*It can be used in areas where only hard water is available.*

# Testing for Carbohydrates

There are three classes of carbohydrates, monosaccharides (such as glucose), disaccharides (such as sucrose), and polysaccharides (such as starch). There are several tests to determine the presence and concentration of each specific class of carbohydrate. It is also possible to test for the presence of carbohydrates generally, without identifying the class. In this experiment you will test for the presence of carbohydrates in general. You will also test for the presence of reducing sugars, sugars that contain free aldehyde or ketone groups.

## PROBLEM

To test for the general presence of carbohydrates. To test for the presence of reducing sugars.

● Part B of this experiment depends upon student familiarity with the structure of aldehydes and ketones. You may want to make this a general experiment for identifying carbohydrates, with less emphasis on identifying reducing sugars.

### Materials

| | |
|---|---|
| Molisch reagent | $H_2SO_4$ (concentrated) |
| Benedict's solution | Bunsen burner |
| 1% glucose | Test tubes |
| 1% sucrose | Beaker |
| 1% starch | |

● Molisch reagent is made by dissolving 10 g of alphanapthol in 100 mL of 95% ethyl alcohol.

## PROCEDURE

### Part A: Molisch Test For Carbohydrates

This is a general test for the presence of carbohydrates.

1. Add 2 drops of Molisch reagent to 2 mL of 1% glucose solution in a test tube. Mix the contents by drumming on the bottom of the tube with your fingers.
2. Holding the tube on a slant, slowly and carefully add 3 mL of $H_2SO_4$ down the side of the tube.
3. Look for the appearance of a blue ring between the two layers. This is a positive test for carbohydrate. Record your observations.

*Positive*

4. Repeat the test using the sucrose solution. Record your observations.

*Positive*

5. Repeat the test, using the starch solution. Record your observations.

*Positive*

● Benedict's solution can be prepared in the following manner. With the aid of heat, dissolve 173 g of sodium citrate and 100 g of anhydrous sodium carbonate in about 800 mL of water. Filter the solution. While stirring slowly, add 17.3 g of copper sulfate dissolved in 100 mL of water to the solution, and then dilute the solution to 1 L. Benedict's solution may also be purchased from any biological supply company.

### Part B: Benedict's Test for Reducing Sugars

This test is specific for reducing sugars.

1. Add 5 mL of Benedict's solution to each of three test tubes.
2. Add 8 drops of the solutions to be tested to the tubes.

3. Place the tubes in a water bath and heat for 3 to 5 minutes. Cool the tubes and compare them. A red color is characteristic of a reducing sugar. Record your observations.

*The tube containing glucose should give a positive result.*

## QUESTIONS

1. Which solutions were positive for the presence of carbohydrates?

*All of them*

2. Which solutions were positive for the presence of a reducing sugar?

*Glucose*

# Tests for Protein

There are several tests for detecting the presence of protein. Since there are many different types of proteins, most of the tests are specific for certain types of proteins or protein products. In this experiment you will use two general tests to detect protein. Most types of protein will give positive results with these tests. The first test you will use is the xanthoproteic test. This work comes from the Greek word *xantho*, which means yellow. The production of a yellow color indicates the presence of protein. The color results when amino or hydroxyl groups on a benzene ring are nitrated. You will also use the biuret test. This is the most general test for detecting protein.

## PROBLEM

To test for the presence of protein.

● Encourage students to bring samples from home to be tested for the presence of protein.

### Materials

10% NaOH
$HNO_3$ (concentrated)
0.1% $CuSO_4$
Test tubes
Bunsen burner
Droppers
Samples of protein
(coagulated albumin, gelatin, preserved meat, fish, milk, rice, and so on)

● Most of the solid samples should be put into solution or at least should be in a small amount of water. Some samples, such as hair or fingernail, can be tested as dry solids.

## PROCEDURE

### Part A: Xanthoproteic Test

1. Place 2 mL (or a small piece) of the sample to be tested in a test tube.
2. Add 1 mL of $HNO_3$. Be careful.
3. Mix the two by swirling the tube and look for a white precipitate.
4. Gently warm the tube. Note any color change. Record the color.

● Caution students about using nitric acid.

● A yellow color indicates protein.

---

5. Allow the mixture to cool. Hold the tube under a stream of tap water after it has cooled off a bit.
6. When the tube is cool add, drop-by-drop, a 10% solution of NaOH. A deep orange color confirms the presence of protein. Record your results in Table 58-1.
7. Repeat the test for the other samples assigned to you.

● Notice that this is a confirmation of the previous color reaction for protein.

### Part B: Biuret Test

1. Place 2 mL (or a small piece) of the sample to be tested in a test tube.
2. Add 2 mL of 10% NaOH solution.
3. Add one drop of 0.1% $CuSO_4$ solution.
4. Mix thoroughly and check for a color change. If there is no color change, add up to 10 drops of $CuSO_4$. Record your results in Table 58-2.
5. Repeat the test for the other samples assigned to you.

● A pink or blue-violet color indicates protein.

**TABLE 58-1 • Xanthoproteic Test**

| SAMPLE | RESULTS |
|---|---|
| | |
| | |
| | |
| | |
| | |
| | |

**TABLE 58-2 • Biuret Test**

| SAMPLE | RESULTS |
|---|---|
| | |
| | |
| | |
| | |
| | |
| | |
| | |

## QUESTIONS

**1.** What color is a positive indication for protein in the xanthoproteic test?

*Yellow (or orange after NaOH is added)*

**2.** What color is a positive indication for protein in the biuret test?

*Pink or blue-violet*

**3.** Which substances tested contained protein?

*Answers may vary.*

# Chemistry of the Fried Egg

Everyone is familiar with the change that occurs when an egg is cooked. It solidifies. But why? Strangely enough, the chemistry involved in the coagulation of protein, such as egg albumin, is an unsolved mystery. The coagulation process apparently involves an unraveling of the long protein molecule, but the exact nature of the process is unknown. It is known that there are several agents that can cause coagulation of protein. We are most familiar with the effect of heat on protein. When an egg is heated, the liquid albumin becomes cloudy, then white and solid. In this experiment you will explore some other ways to coagulate protein.

● Impress upon students that heating is only one of many ways to "cook" an egg.

## PROBLEM

To investigate the coagulation of protein

### Materials

● **CAUTION** Both $HgCl_2$ and $Pb(NO_3)_2$ are poisonous.

Fresh egg
2% $AgNO_3$
5% $HgCl_2$
10% tannic acid solution
95% ethyl alcohol
10% $Pb(NO_3)_2$
Droppers
Test tubes
Beaker
Bunsen burner

## PROCEDURE

Prepare an albumin solution for use in all parts of this experiment by adding the white of one egg to 100 mL of water in a beaker. Stir until the albumin is in solution.

### Part A: Effect of Heat on Protein

● You can prepare a class-quantity of albumin from the commercial extract. However, it gives the experiment a more "practical" flavor to use the albumin from eggs.

1. Place 5 mL of the albumin solution in a test tube.
2. Heat the test tube in a boiling water bath for 5 minutes. Describe the results.

*The albumin coagulates.*

### Part B: Effect of Alcohol on Protein

1. Add 5 mL of albumin solution to a test tube.
2. Add 10 mL of ethyl alcohol.
3. Describe the results.

*The albumin coagulates.*

### Part C: Effect of Tannic Acid on Protein

● The effect of tannic acid on protein is the basis for the tanning process for treating animal hides.

1. Add 5 mL of albumin solution to a test tube.
2. Add 2 mL of 10% tannic acid solution.
3. Describe the results.

*The albumin coagulates.*

● An emergency treatment for ingestion of certain toxic substances is to force-feed egg whites. Some of the toxin is absorbed in the process of coagulation of egg protein.

**Part D: Effect of Salts of Heavy Metals on Protein.**

1. Place 5 mL of albumin solution in each of three test tubes.
2. To the first tube add 2% $AgNO_3$, drop-by-drop. Describe the results.

*The albumin coagulates.*

3. To the second tube add 5% $HgCl_2$, drop-by-drop. Describe the results.

*The albumin coagulates.*

4. To the third tube add 10% $Pb(NO_3)_2$, drop-by-drop. Describe the results.

*The albumin coagulates.*

## QUESTIONS

1. Why does a doctor wipe the skin with alcohol before giving an injection?

*The alcohol kills the bacteria by coagulating the bacterial protein.*

2. People who have accidentally swallowed Hg, Ag, or Pb salts may be given egg white. Why?

*The metal ions are tied up in the coagulated albumin.*

## EXPERIMENT 60

Chapter 20

# Are Your French Fries Greasy?

● This experiment involves the use of organic solvents. Great care should be used. The extraction must be carried out under a hood or in a well-ventilated area.

You have probably noticed that paper in which meat has been wrapped often has grease spots. You might obtain the same effect if you dropped food on your homework paper. The characteristic spot that fat or grease leaves on clear paper is a simple test for the presence of fat. In this experiment you will extract the fat from a food sample by dissolving it in a fat-soluble substance. You will also extract fat from several samples of french fried potatoes and compare the amount of fat in the samples.

## PROBLEM

To test for the presence of fat in a sample of food. To compare the amount of fat in several samples of french fries.

### Materials

Freon
Unglazed paper
Beakers
Flasks
Filter paper
Droppers

● It is *much more advisable* to use freon than carbon tetrachloride. Freon may be purchased from anyone who services refrigerators or air conditioners. Carbon tetrachloride can be used as a last resort.

## PROCEDURE

● Have several samples of french fries from nearby fast food restaurants available. You might want to have the students bring in the samples.

### Part A: Testing for Fat

1. Place 1 or 2 french fries in a 250-mL flask.
2. Add 25 mL of freon to the flask.
3. Swirl the flask for a few minutes to dissolve the fat in the solvent.
4. Using a dropper, place a drop or two of the solution in the center of a sheet of unglazed paper.
5. Allow 5 minutes for the solvent to evaporate.
6. Hold the paper up to the light and look for a translucent spot. Record your observations.

*A grease spot appears.*

### Part B: Extracting Fat

1. Select 4 or 5 french fries for your sample. Determine the mass of the sample. Record the mass in Table 60-1.
2. Place the french fries in a 250-mL flask.
3. Add enough freon to cover the french fries.
4. Shake the flask for a couple of minutes to dissolve the fat.
5. Carefully pour off the solvent. (If you pour the solvent down the drain, flush with water. Be careful not to lose any of the sample.)
6. Repeat steps 3 through 5 twice more.

7. Remove the potatoes from the flask. Put them on a piece of filter paper.
8. Place the potatoes under a hood, or in a well-ventilated area and allow the solvent to evaporate.
9. Determine the mass of the greaseless potatoes. Record the mass in Table 60-1.
10. Complete Table 60-1.
11. Repeat the procedure for each sample assigned to you.

**TABLE 60-1 • Fat in French Fries**

| | SAMPLE 1 | SAMPLE 2 | SAMPLE 3 |
|---|---|---|---|
| Mass of French Fries | | | |
| Mass of French Fries after Solvent Evaporates | | | |
| Mass of Fat | | | |
| % Fat | | | |

## QUESTIONS

1. Which sample of french fries contained the most fat?

2. What is the purpose of adding freon to the samples?

*The fat dissolves in the freon.*

EXPERIMENT 61

Chapter 20

# The Fat Content of Milk

● This experiment is intended to give students experience in quantitative techniques of separation. Use extreme care with the solvents. All work must be done under a fume hood, or in a well-ventilated area. Stress the simplicity of the mathematics involved in calculating percent.

In Experiment 60 you determined the amount of fat in several samples of french fries. In this experiment you will determine the amount of fat in a sample of milk. A solvent will be used to extract the fat. As in the previous experiment, you will determine the mass of the sample before and after the fat is extracted. As a check on your results, you will allow the solvent to evaporate. You can then determine the mass of the residue. This represents the fat extracted by the solvent.

## PROBLEM

To calculate the percent of fat in milk.

### Materials

● Other types of milk may be tested.

Fresh, whole milk
Chloroform
Food coloring
Steam bath
Separatory funnel
Ringstand and ring
Evaporating dish
Graduated cylinder
Beakers

## PROCEDURE

1. Determine the mass of a 50-mL sample of whole milk. Record the mass in Table 61-1.
2. Place the milk in a separatory funnel.
3. Add 30 to 50 mL of chloroform to the funnel. Be careful not to breath the chloroform fumes.
4. Add 2 drops of food coloring to the funnel.
5. With the stopper in place, shake the two liquids for a couple of minutes. Place the funnel in a ring in the ringstand.
6. Determine the mass of an evaporating dish. Record the mass in Table 61-1.
7. Draw off the chloroform layer (lower layer in the funnel) into the evaporating dish.
8. Place the evaporating dish in a steam bath.
9. Draw off the remaining milk into a beaker and determine the mass. Record the mass.
10. After all of the chloroform has evaporated, determine the mass of the evaporating dish plus the residue. Record the data you obtain in Table 61-1.
11. Complete Table 61-1.

● Have the students use a 150-mL beaker when determining the mass of the milk sample. The mass of the beaker can be determined beforehand or the adjustable mass on the balance can be set to compensate for the beaker's mass.

● Students will have to observe carefully to distinquish the two layers.

● If a steam bath is not available, the dish may be placed on top of a beaker of boiling water.

● Most of the caloric value of milk is in the fat. One cup of milk represents about 160 kcal.

**TABLE 61-1 • Fat in Milk**

| | |
|---|---|
| Mass of Milk | |
| Mass of Evaporating Dish | |
| Mass of Milk after Fat Extraction | |
| Mass of Evaporating Dish plus Residue | |
| Mass of Milk (Fat) Lost | |
| Mass of Residue | |

● Results may range from 1% to 10%. Average values are:
Whole milk—4.5-5%
Condensed milk—5-7%
Dry milk—8%

## QUESTIONS

1. Determine the percent of fat in the milk sample from the mass of the milk sample before and after the fat was extracted.

2. Determine the percent of fat in the milk sample from the mass of the sample and the mass of the residue.

# The Fat Content of Ground Beef

In this experiment you will make a rough determination of the percent of fat in several samples of ground beef. The amount of fat permitted in ground beef is regulated by law: 30% maximum for ground beef and 15% maximum for lean ground beef. You will use a 100-g sample of ground beef. You will boil off and collect the fat and measure its volume directly in a graduated cylinder. Since 1 mL of fat has a mass of approximately 1 g, you can read the percent of fat directly from the graduated cylinder.

● A little consumer chemistry can be done here. Students should be encouraged to bring in samples of ground beef from various supermarkets in town. Have students record the price per pound. This can then be correlated with the percent of fat.

## PROBLEM

To determine the percent of fat in ground beef.

### Materials

| | |
|---|---|
| Beakers | Samples of ground beef and lean ground beef |
| Ringstand | Graduated cylinder (100 mL) |
| Bunsen burner | Stirring rod |
| Tongs | |

● Have large-sized filter paper available for determining the mass of the beef. Instruct the students to compensate for the mass of the paper.

## PROCEDURE

1. Using a balance, obtain a 100-g sample of ground beef. (The mass of the sample should be as close to 100 g as possible.) Place the sample in an 800-mL beaker.
2. Add 600 mL of water to the beaker. Stir thoroughly and bring to a boil.
3. Gently boil the sample for 15 minutes. Be careful not to allow any material to spill or spatter.
4. Remove the beaker from the heat. Allow the meat to settle to the bottom. The fat will form a liquid layer on top of the water.
5. Carefully pour off the fat layer into a clean, 100-mL graduated cylinder. It may be difficult to remove the last traces of fat. When the beef and water mixture cools, pick out the solidified fat particles. Add these to the graduated cylinder.
6. Read the volume of fat in the cylinder. You are assuming that 1.0 mL is roughly equivalent to 1.0 g of fat. Since your original sample of meat was 100 g, the volume is also the percent of fat. Record your data in Table 62-1.
7. Repeat the procedure using the other samples assigned to you.

● Since this is about ¼ pound, you might want students to work in groups of 4 or 5.

● If pistol-grip tongs are not available, turn off the burner and allow the beaker to cool.

● Remind students that this is what they commonly see when leftover chili, beef stew, and so on, is placed in the refrigerator—the fat separates and comes to the top.

**TABLE 62-1 • Fat in Ground Beef**

● Emphasis here is on percent determination.

| | SAMPLE 1 | SAMPLE 2 | SAMPLE 3 |
|---|---|---|---|
| Mass of Ground Beef | 100 g | 100 g | 100 g |
| Volume of Fat | | | |
| Percent Fat | | | |

## QUESTIONS

● Chart this on the chalkboard, using all students' data.

1. What is the relationship between the price per pound of the ground beef and the fat content?

*Answers will vary.*

2. How does heat affect the fat in the ground beef?

*Heat liquefies the fat so that it can be separated.*

# Testing for Starch

Starch is a polysaccharide consisting of a long chain of glucose molecules. Upon hydrolysis, starch is first reduced to dextrins and then to molecules of the disaccharide maltose. Heat will speed up the hydrolysis process, quickly reducing starch to maltose. Starch reacts with iodine to produce an intense, blue-black color. The nature of the starch-iodine complex is not known. However, it is suspected that the coiled and twisted chain of glucose molecules traps the iodine, producing the colored complex. Dextrins give a reddish color with iodine. Maltose gives no color with iodine. Presumably, hydrolysis destroys the coiled chain and thus prevents the colored complex from forming.

## PROBLEM

To test for starch. To follow the hydrolysis of starch.

### Materials

Starch solution
Iodine solution
Droppers
Test tubes
Foods containing starch
(potatoes, bread, crackers)
Bunsen burner
Test-tube holder

● Make up a starch solution for class use. Be sure to use soluble starch!

## PROCEDURE

1. Add 1 mL of the starch solution to 5 mL of distilled water in a test tube.
2. Add a few drops of iodine solution. Record your observations.

*A blue-black color appears.*

● Iodine solution is made by adding a few grains of $I_2$ to a small amount of KI in alcohol.

3. Carefully heat the solution to boiling. Record your observations.

*The color fades.*

4. Set the tube aside and allow it to cool. Record your observations.

*A red color appears.*

5. Test each food sample assigned to you for the presence of starch by adding a drop of iodine to the sample. Record your observations.

## QUESTIONS

1. Explain your observations for step 3.

*Heating breaks down the starch into maltose molecules, thus no color is evident.*

2. Explain your observations for step 4.

*On cooling, dextrins form, combine with iodine, and produce the red color.*

# Starch in Plant Leaves

Through the process of photosynthesis, green plants convert carbon dioxide and water into sugar. Since this is an endothermic reaction, energy must be provided. The energy comes from sunlight. Solar energy is captured by chlorophyll, the green pigment in plants. This energy is used to produce the sugar. The plant converts much of the sugar into starch. The starch is stored by the plant until it is needed.

## PROBLEM

To test for the presence of starch in a green leaf.

### Materials

Ethyl alcohol
Iodine solution
Green leaves
Beakers
Bunsen burner
Ringstand
Forceps
Tongs
Test tube

● Obviously, this experiment must be done during the spring or early fall when there are *still* green leaves around.

## PROCEDURE

1. Select 2 green leaves. Tear them into several pieces and place them in a 250-mL beaker.
2. Add 100 mL of water to the beaker. Boil the leaves for 5 minutes.
3. Turn off the burner. Remove the beaker from the ringstand. (Use tongs!)
4. Fill a test tube ½ full of ethyl alcohol.
5. Using forceps, remove a piece of leaf from the beaker and place it in the test tube.
6. Place the tube in a beaker of water. Gently heat the water for five minutes.
7. Remove the leaf from the test tube. Rinse the leaf with water and blot it with a paper towel.
8. Place the leaf on a paper towel and add a few drops of iodine. Record your observations.

● **CAUTION** Warn the students to be careful when heating the test tube in the beaker of water. Alcohol is flammable. Use a hotplate for this step if one is available.

*The leaf turns blue-black.*

## QUESTIONS

1. Why is it necessary to heat the leaf before testing it with iodine?

*Heating dissolves the wax coating on the leaf, allowing the starch to be tested.*

2. What evidence is there that the green leaf contained starch?

*The blue-black color with iodine indicates starch.*

# Personal Care Chemistry

Cosmetic and drug companies spend millions of dollars annually on research to improve their products. These companies employ thousands of chemists. In Experiment 39 you made cleansing cream that was almost identical to the cleansing creams sold in drugstores. In this experiment you will prepare two more personal care items. The first is underarm deodorant. The second is athlete's foot powder.

## PROBLEM

To prepare personal care items by duplicating the procedures used in the cosmetic industry.

● Some students might be allergic to these chemicals when used in personal-care items. It is advisable to try a small amount first.

### Materials

Isopropyl alcohol
Zinc oxide
Boric acid
Talc
Sodium thiosulfate
Plastic squeeze bottle
Alum (aluminum potassium sulfate)
Mortar and pestle
Inexpensive saltshaker
Beakers
Wooden splints

## PROCEDURE

### Part A: Preparing Deodorant Liquid

1. Add 65 mL of isopropyl alcohol to 150 mL of distilled water in a 400-mL beaker.
2. Place 1.5 g of alum and 1.5 g of zinc oxide in a clean, dry mortar. Using the pestle, grind the two solids to a fine powder.
3. Dissolve the powdered mixture in the alcohol solution in the beaker.
4. Place the deodorant liquid in the squeeze bottle. To use, shake the container, squirt the liquid onto cotton applicators and rub lightly under the arms. (Use sparingly!)

● Be sure to shake the container, since the alum and zinc oxide tend to settle out.

### Part B: Preparing Athlete's Foot Powder

1. Using a mortar and pestle, grind 15 g of boric acid into a fine powder. Place this in a small beaker.
2. Clean and dry the mortar, then grind 5 g of sodium thiosulfate into a powder. Add this to the boric acid in the beaker.
3. Dry talc is usually in the form of a fine powder. If your talc is coarse, grind it as you did the other two solids. Add 10 g of the talc to the beaker.
4. Mix the three solids thoroughly with a wooden splint.
5. Place the mixture in the saltshaker. Sprinkle it on the feet while dressing in the morning.

● An inexpensive saltshaker is a good dispenser for this mixture of powders.

# Chemistry and Crime Detection: Blood

Police department personnel and criminal investigators often need assistance from chemists to help solve crimes. One common procedure is to determine if certain stains are dried bloodstains. In this experiment you will use three tests to detect the presence of blood in a stain.

## PROBLEM

To determine if a stain contains dried blood.

### Materials

Samples of stained paper and fabric
Test tubes
0.2% acetic acid
Graduated cylinder
Droppers
3% $H_2O_2$
Benzidine solution
Phenolphthalein
Beaker
2 *M* NaOH
Zinc dust
Ringstand
Bunsen burner
Brown bottle with cap
Mossy zinc
20% $H_2O_2$
Evaporating dish
$H_2SO_4$ (concentrated)
Ultraviolet lamp
Scissors

● Most clinics or the local blood bank will give you a small container of old blood. Prepare samples of cloth and paper stained with catsup, food coloring, and so on. Have the students detect the real blood sample.

● The bottle should be about 500 mL in size.

## PROCEDURE

### Part A: Benzidine Test

This test is based upon the ability of the hemoglobin in blood to rapidly oxidize benzedrine in the presence of $H_2O_2$. The appearance of a blue color indicates the presence of blood.

1. Cut out a small piece of the paper or fabric containing the stain.
2. Place it in a test tube and add 5 mL of 0.2% acetic acid.
3. Add 5 drops of 3% $H_2O_2$.
4. Add 2 mL of benzidine solution.
5. If the stain is blood, a blue (or blue-black) color will form. The color will eventually change to red. Record your observations.

_____

_____

### Part B: Phenolphthalein Test

To perform this test, you must first prepare a test solution as follows. In a beaker dissolve 1 g of phenolphthalein in 100 mL of NaOH. Add 10 g of zinc dust. Warm the solution until it becomes colorless. Place a small piece of mossy zinc in a brown bottle. Add the solution to the bottle.

1. Add 5 mL of the test solution to a test tube.
2. Add 5 mL of 20% $H_2O_2$.

3. Add a small sample of the material containing the stain.
4. If the stain is blood, a red color will result. Record your observations.

_____

_____

● **CAUTION** Students must be warned about damage to the eyes that could result from looking into the UV lamp.

**Part C: Ultraviolet Test**

Like many other substances, blood will fluoresce, giving a characteristic color, when exposed to ultraviolet light.

1. Place a sample of the stained material in an evaporating dish.
2. Add 3 to 4 drops of concentrated $H_2SO_4$ to the stain. Caution!
3. Place the dish under an ultraviolet lamp.
   **(CAUTION: DO NOT LOOK DIRECTLY INTO THE UV LAMP!)**
4. A scarlet fluorescence is characteristic of blood. Record your observations.

_____

_____

Repeat all of the tests with samples supplied by your teacher. Record your observations.

_____

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EXPERIMENT 67 Exc. 13

# Chemistry and Crime Detection: Paint and Forgery

In Experiment 66 you saw how chemistry can be useful in solving crimes by detecting blood. Chemists are involved in many other identification techniques in police work. In this experiment you will perform a simple test to match paint samples. You will also test to determine if a "document" has been forged.

## PROBLEM

To match samples of paint. To identify a forged signature.

### Materials

Samples of wood with paint stains
Ultraviolet lamp
Sample documents with signatures

## PROCEDURE

### Part A: Matching Paint Samples

Like blood, paint fluoresces under ultraviolet light. The colors produced are so characteristic that only samples from the same batch of paint will give the same colors.

● Obtain two brands of the same shade of green paint. Paint two pieces of wood with one brand and a third piece with the other brand. Label the samples 1, 2, and 3. (Keep a key!)

1. Obtain samples of wood with paint stains from your teacher.
2. Place the samples under an ultraviolet lamp.
3. Determine which two samples contain paint from the same can. Record your observations.

4. Since paint is almost always available in hit-and-run cases, describe how this test would be useful.

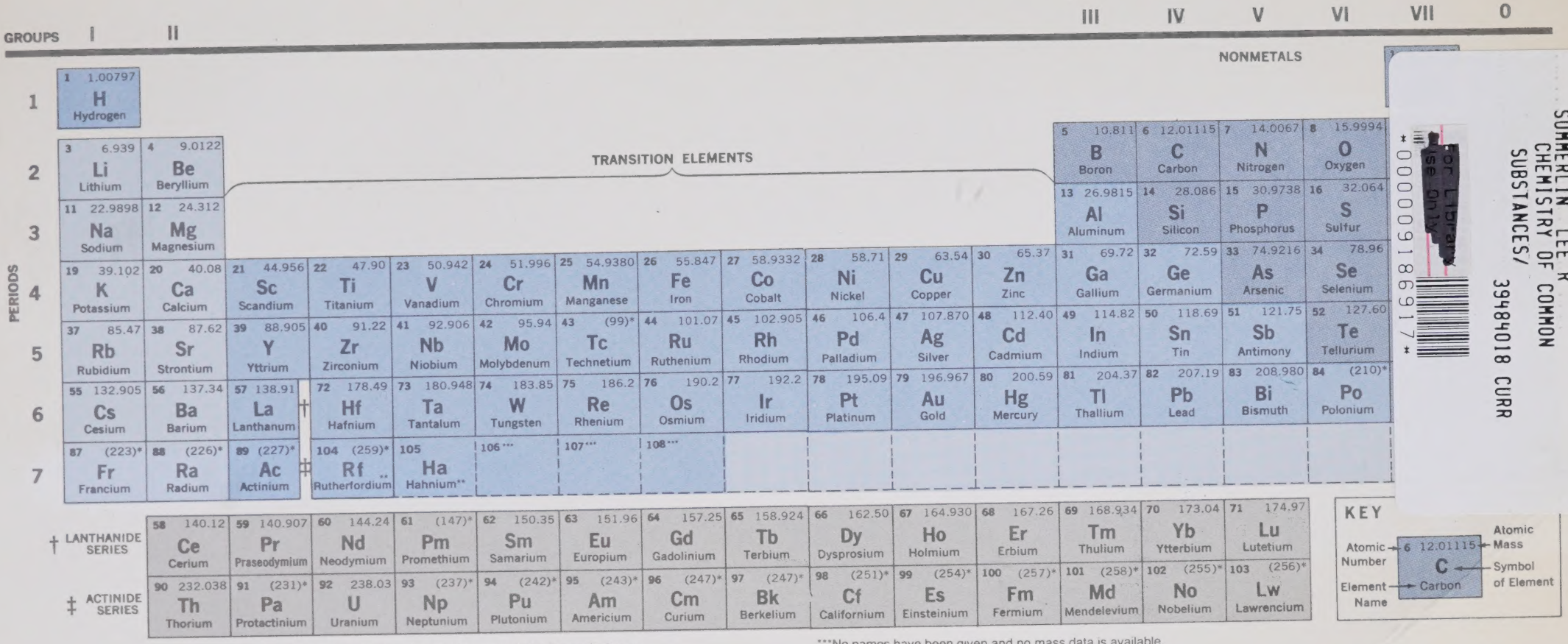
*The auto involved in the case can be identified by matching a paint sample from the crime with the paint on the auto.*

● Prepare three short "documents" and sign them with the name of an historical person. Fountain pen must be used. Bleach off the signature from one document. Have another teacher "forge" the historical signature by copying your writing. Label the documents 1, 2, and 3. (Old, cancelled checks can be used.)

**Part B: Detecting Forged Documents**

When documents or checks are forged, quite often the original signature is erased by using a bleaching agent. However, the forger seldom bothers to rinse away the residual ink or bleach. These will easily show up under an ultraviolet lamp.

1. Obtain the sample documents from your teacher.
2. Place the documents under the ultraviolet lamp.
3. Determine which document has a "forged" signature. Record your observations.




*Atomic masses appearing in parentheses are those of the most stable known isotopes.

**Names are unofficial.

***No names have been given and no mass data is available.

Atomic masses based on C-12 = 12.0000

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CHEMISTRY OF COMMON
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# Chemistry of Common Substances

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